Elements of Physical Science

Das Gupta and Banerjee

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Recommended by the West Bengal Board of Secondary
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ELEMENTS OF PHYSICAL SCIENCE

[FOR CLASS IX]

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PREFACE

From 1974, eleven-year higher secondary course will be replaced by ten-year secondary course. The Board of Secondary Education has already circulated the syllabi of different subjects included in the new course. One of the salient features of this new course is the emphasis on science which has been made a subject of compulsory study for all boys and girls. Two papers—one on physical science and the other on life science—have been set apart for students going in for school final examination. This book has been written according to the syllabi of physical science (Physics and Chemistry) for Class IX.

Attemps have been made to present the subject matter to the young learners in a lucid and simple manner. Science-teaching at the ordinary level should be, in our opinion, based mainly on experiments and demonstrations. We have, therefore, tried to explain the elementary facts of science through simple experiments, wherever possible.

We shall consider our labours amply rewarded if the book is found suitable by the young readers for whom it is meant.

Calcutta
24th. Dec., 1973

C. R. Das Gupta
S. P. Banerjee

SYLLABUS

CLASS IX

Common to both Physics & Chemistry

(A single text-book covering the items)

1. Systems of measurement. Different physical quantities and their units. Measuring devices : Ordinary scale, common balance, measuring cylinders, clocks. [14 Pages]
2. Matter and energy. Mass and weight, conservation of mass and conservation of energy. Different forms of energy. Transformation of energy (non-quantitative treatment). [8 Pages]
3. Change of state. Freezing, melting, boiling, evaporation and condensation. Melting point and boiling point. Factors effecting them. Idea of latent heat. [8 Pages]

PHYSICS

1. Rest and motion ; Displacement, speed, velocity, acceleration and retardation. Newton's laws of motion. Definition of force. (Equations of motion excluded . Rotational motion excluded) [10 Pages]
2. Work, energy and power—potential and kinetic energy. Units (No sums). Simple machines ; inclined plane, wheel and axles and levers (Non-mathematical). [9 Pages]
3. Nature of heat. Heat and temperature. Factor determining the quantities of heat. Heat as a form of energy. Relationship with work. [9 Pages]
4. Source of light. Ray of light. Reflection and Refraction of light. Total internal reflection. Explanation of some natural phenomena on the basis of refraction and total reflection. Propagation and velocity of light. Convex lens and its focussing action. Focal length. Convex lens as a magnifying glass. Dispersion of light. Spectrum (Demonstration). [20 Pages]

CHEMISTRY

1. (i) Physical states of matter—reasons for existence of the three states—Melting and boiling points.

   (ii) Identification of matter—physical and chemical properties—How matter (solid, liquid and gas) differs in physical properties (touch, colour, smell, solubility, magnetic, property, etc). and chemical properties (behaviour on heating, treatment with acids, etc)

   (iii) Physical and chemical changes : factors which induce and regulate chemical change, viz., contact, temperature pressure, catalysis, etc. Exothermic and Endothermic change.

   (iv) Elements and compounds—Metals and Non-metals. [16 Pages]
2. Solution—Solvent and solute : Unsaturated and saturated solution. Solubility and its relation with temperature. [4 Pages]
3. Symbols, formulae and chemical equations : Significance of chemical equations and balancing simple equations. [5 Pages]
4. Electrolysis—Electrolytes and non-electrolytes—Decomposition of water : Electroplating. [6 Pages]
5. Ideas of acids, bases and salts : Neutralization. [3 Pages]
6. Oxidation and reduction. [2 Pages]
7. Liquid air. Nitrogen and CO_2 cycle (Elementary ideas) : use of the rare gases in air, Neon lighting. [5 Pages]
8. Simple method of preparation and properties of Oxygen, Hydrogen, Nitrogen, Ammonia, Carbon dioxide, Sulphur dioxides, Sulphuretted hydrogen. [12 Pages]

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COMMON TOPICS

CHAPTER ONE

Units and Systems of Measurement

1·1. Physical quantity :

Anything which can be measured is a quantity. For example, we know that a block of wood has some weight and we can measure that weight by a balance. Hence, '*weight*' is a quantity. If any event takes place for a certain time, we can measure the time by a clock. So, '*time*' is a quantity. Similarly, the length, breadth or height of a room may be measured either by a scale or by a tape. *Length, breadth* or *height*, are therefore, quantities. In the same manner, *volume*, *mass*, *velocity*, *acceleration*, *electric current strength*, etc. are all examples of different quantities because each one of them may be measured by suitable instruments. In studying Physics, we shall come across numerous such quantities. All such quantities included in Physical Science, are called physical quantities.

Definition : *Anything which is measurable is a physical quantity.*

1·2. Uses of measurement :

In order to express the measure of a quantity, we generally choose a convenient amount of the quantity as standard and express the measure of the quantity or other similar quantities in terms of that standard. This standard of measurement is called a *Unit*. Illustrations will make the idea clear.

Suppose, we say that a room is 20 feet long, From this statement we find no difficulty in realising how long the room is. Here, *length* is a quantity and for the measure of the length, '*foot*' has been taken as its unit.

If I say that I have purchased a lot of rice, then I do not give exact idea about the quantity of rice purchased. But if I say that I have purchased 20 kilograms of rice, then one can, at once, form an exact idea regarding the quantity of rice. Here, '*kilogram*' has been used as a unit to express the measure of the *mass* of rice.

In the same way, if it is said that the Bombay Mail took a long time to reach Calcutta, then time is not mentioned exactly. To give an exact impression of the time taken, it should be mentioned 30 hours or 40 hours or so. Here, *hour* has been used at the unit to measure *time*.

In this way, it is found that each quantity requires a unit for its measurement. But there are thousands of quantities in Physics. Do they have thousands of units ? Fortunately it is found that although the quantities are numerous in numbers, the units are not. If we fix up units of only three quantities, the units of all other quantities may be formed from them. These three quantities are (i) **Length**, (ii) **Mass** and (iii) **Time**. The units of these three quantities do not depend upon one another and they are known as **Fundamental Units**. Units of other quantities, which can be formed from one or more of the fundamental units are called *Derived Units*.

Definition : *The units of those three quantities, which are independent of one another and from which the units of other quantities may be derived, are called fundamental units.*

Those units which are derivable from one or more of the fundamental units, are known as derived units.

1·3. Different systems of units :

There are, in general, two systems to express the above mentioned three fundamental units.

(1) *C.G.S. or French or Metric system :*

The word 'C' means Centimetre →unit of length.
,, ,, 'G' ,, Gram →unit of mass.
,, ,, 'S' ,, Second →unit of time.

(2) *F.P.S. or British system :*

The word 'F' means Foot →unit of length.
,, ,, 'P' ,, Pound →unit of mass.
,, ,, 'S' ,, Second→unit of time.

This system is widely used in British empire.

Apart from these two systems, there is a third system known as M. K. S. system. According to this system, the units of length and mass are a metre (M) and a kilogram (K) respectively while that of time is a second (S).

1·4. Units of length :

Centimetre : It is the unit of length in the C.G.S. system.

The distance, at 0°C between two lines on a platinum-iridium bar kept at the International Office of Weights and Measures in France, is taken as a *Metre*. Centimetre is one-hundredth part of a metre. Multiples and sub-multiples of metre have been made to measure small and large lengths. One of the advantages of the C.G.S. system is that these units of small and large lengths are related to the metre by either multiples or sub-multiples of 10 as will be evident from the table below :

10 Millimetres (m.m.) = 1 Centimetre (cm.)
10 Centimetres = 1 Decimetre
10 Decimetres = 1 Metre (m.)
10 Metres = 1 Decametre
10 Decametres = 1 Hectometre
10 Hectometres = 1 Kilometre (Km)

Foot : It is the unit of length in the F.P.S. system.

The distance, at 62° Fahrenheit, between two lines on a bronze bar kept at the British exchequer at London, is taken as a *Yard*. Foot is one-third of a yard. Multiples and sub-multiples of foot, for the measurement of large and small lengths, are as follows :

1 mile = 1760 yards.
1 yard = 3 feet.
1 feet = 12 inches.

Remember, 1 inch = 2·54 centimetres.
1 metre = 1·09363 yards = 39·37 inches.
1 kilometre = 0·621 mile.

1·5. Units of Volume :

If the length, breadth and height of a cube, is each 1 cm., then the volume of the cube is taken as the unit of volume in the C.G.S. system (Fig. 1). It is called *one cubic centimetre* or one *c.c.*

Likewise, if the length, breadth and height of a cube, is 1 foot each, then its volume is taken as the unit of volume in the F.P.S. system and it is called 1 *cubic foot* or 1 *cu. ft.*

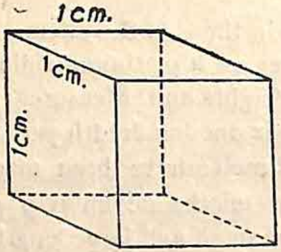

Fig. 1

In the C.G.S. system, *litre* is also used as a unit of volume—specially in the case of liquids.

1 litre=1000 cubic centimetres.

Similarly, in the F.P.S. system '*Gallon*' is used as a unit to express the volume of liquids.

1 gallon=volume of 10 pounds of water at 60°F
=0·16 cubic foot.

Again, 1 cu. ft.=6·25 gallons.

Remember, 1 gallon=4·54 litres.

To measure the volume of liquids, a measuring cylinder, graduated in cubic centimetres (c.c.) is usually used. Fig. 2(a) shows such a measuring cylinder. Readings on this instrument are always taken at the level of the bottom of the curved surface of the liquid, avoiding parallax as shown in Fig. 2(b). The figure

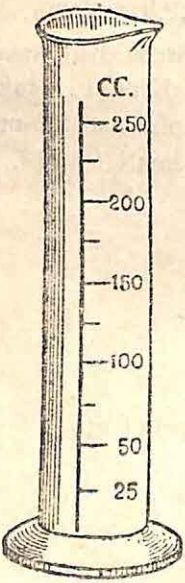

Fig. 2(a)

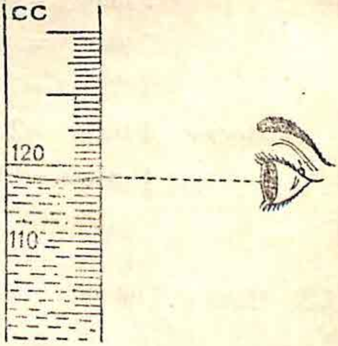

Fig. 2(b)

shows the correct position of the eye and the reading corresponding to that position is 118·5 c.c. Hence, the correct volume of the liquid inside the cylinder is 118·5 c.c.

1·6. Units of mass :

Definition : *Mass of a body means the quantity of matter contained in the body.* For example, mass of an iron ball is the amount of iron contained in the ball.

Gram : It is the unit of a mass in the C. G. S. system. The mass of a certain cylindrical piece of platinum-iridium alloy kept in Paris is called a standard *Kilogram,* which is known as the International Prototype Kilogram. Gram is one thousandth part of a Kilogram.

Usually the mass of 1 c.c. of water at 4°C is taken as 1 gram. Following are the multiples and sub-multiples of a gram :

10 miligrams (mgm) = 1 centigram
10 centigrams = 1 decigram
10 decigrams = 1 gram (gm)
10 grams = 1 decagram
10 decagrams = 1 hectogram
10 hectograms = 1 kilogram (Kgm)
100 kilograms = 1 quintal.

Pound : It is the unit of mass in the F. P. S. system and is abbreviated as lb.

The mass of a certain piece of platinum kept at the Standards department of the Board of Trade in England is taken as a pound and is known as the Prototype Pound. Other units of mass, in use, in the F. P. S. system are as follows :

16 drums = 1 ounce (oz)
16 ounces = 1 pound (lb)
28 pounds = 1 quarter
4 quarters = 1 hundred-weight (cwt)
20 hundred-weights = 1 ton
So, 1 ton $= 20 \times 4 \times 28 = 2240$ pounds.

Remember, 1 pound = 453·59 grams = 0·4536 kilograms.
1 kilogram = 2·204 pounds.

1·7. Units of Density :

Up to this point you have become acquainted with units for measuring length and mass. These are called fundamental units because the great variety of other quantities that we meet in Physics can be expressed as combinations of them. We have already had one such derived unit—the unit of volume which is based on simple combination of length unit.

As a further example, let us have a look at a useful quantity called *density*. Everybody realises that a given volume of one material has, in general, a different weight than the same volume of some other material. For instance, we ordinarily say that iron is 'heavier' than wood. More exactly we should say that *any given volume* of iron is heavier than *an equal volume* of wood. To make the comparison exact, we can weigh a certain volume of iron, say 1 c.c. When this is done, the weight is found to be 7·9 gms. By comparison, the weight of 1 c.c. of wood is about 0·5 gms. We say that the density of iron is 7·9 gms/c.c. while that of wood is 0·5 gm/c.c. Thus, density is defined as *mass per unit volume of the substance.*

If the mass of V volume of a substance is M, then its density

$$D=\frac{M}{V} \text{ or, } M=D\times V.$$

It is to be noted that in the C. G. S. system, the unit used for density is gms/c.c. In the F. P. S. system, the unit is ℔/cu. ft.

It is to be further noted that the density of water, in the C. G. S. system is 1 gm/c.c, while in the F. P. S. system, it is 62·5 ℔s/cu. ft.

1·8. Unit of time :

Due to the rotation of the earth about its own axis, the sun appears to travel across the sky from east to west. The interval between two successive crossings of the meridian at a place by the sun is called a *solar day*. For some reasons, the length of a solar day instead of remaining constant throughout the year changes slowly. The average taken over a year is known as *mean solar day.* $\frac{1}{24}$th part of a mean solar day is an hour. $\frac{1}{60}$th part of an hour is a minute, while $\frac{1}{60}$th part of a minute is called a second. So,

24 hours=1 mean solar day ; 60 minutes=1 hour

60 seconds=1 minute or,

$$1 \text{ second}=\frac{1}{60\times60\times24}=\frac{1}{86400}\textit{th} \text{ part of a mean solar day.}$$

Definition : In both the systems—C. G. S. and F. P. S. —the unit of time is the mean solar second or 'second' and it is $\frac{1}{86400}$*th* part of a mean solar day.

1·9. Advantages of Metric or Decimal System :

If you look at the units of length or mass in the metric system, you will see that each unit is ten times its next lower unit or one-tenth of its immediate higher unit. For this reason the metric system is known as decimal system. One great advantage of this system is that for conversion of one unit into the higher or lower unit, shifting of decimal point will suffice. For example, 593·21 metre=59321 centimetre=0·59321 kilometer etc. The F. P. S. system does not possess the advantage. For example, 3 yards$=3\times3=9$ ft$=9\times12$ inches$=\frac{3}{1760}$ mile, etc. Moreover, units of length, volume and mass are conveniently related in the metric system. As for example, mass of 1 cubic centimetre of water is 1 gram. But the mass of 1 cubic foot of water is not 1 ℔—it is 62·5 ℔s.

For this reason, decimal systems have been adopted by most of the countries in the world. In India decimal systems have been adopted in currency from 1957 and in weights from 1961.

Measurement of length, mass and time

1·10. Measurement of length :

We generally use scale for the measurement of length. It is a properly seasoned thin regular bar of box-wood or of metal, graduated in its lower part in centimetres and millimetres and in upper part in inches and fractions thereof [Fig. 3(a)]. Sometimes, a scale is graduated only in centimetres and millimetres in which case, the scale is called a *metre scale.* Likewise, when a scale is graduated only in inches and fractions thereof, the scale is called *foot-rule.*

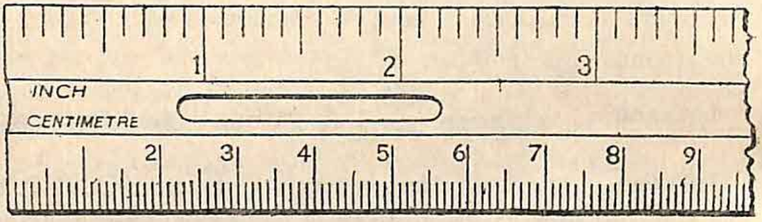

Fig. 3(a)

Use of a scale : Suppose a scale is to be used to measure the length of the line *AB.* For this, the scale is to be placed in such a manner that the graduated side of the scale may be as close as possible to the line *AB* and the end *A* of the line may be coincident with a complete division (say, 1 cm. mark) of the scale. In this postion, you are to take the reading of the other end *B* of the line, looking *perpendicular* to the scale. Suppose, the end *B* lies between 8·8 cm. and 8·9 cm. [Fig. 3(b)]. In such cases, the exact position of the end *B* should be read, on eye-estimation, by mentally dividing 1 mm. into ten equal parts. *B*-end reading is, suppose, 8·89 cm.

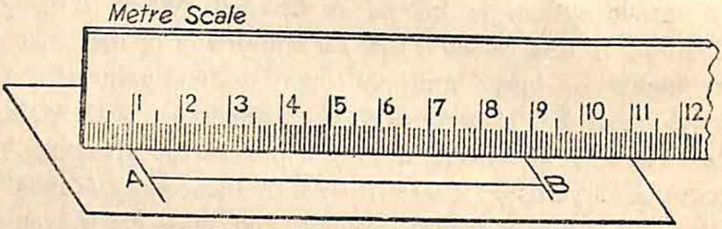

Fig. 3(b)

Hence, the length of the line $AB = B$-end reading$-A$-end reading
$= 8{\cdot}89 - 1 = 7{\cdot}89$ cm.

In this way, you are to take readings several times, using different portions of the scale and the mean of those readings will give you the exact length. In the same way, the length of the line *AB* may be measured in inches also.

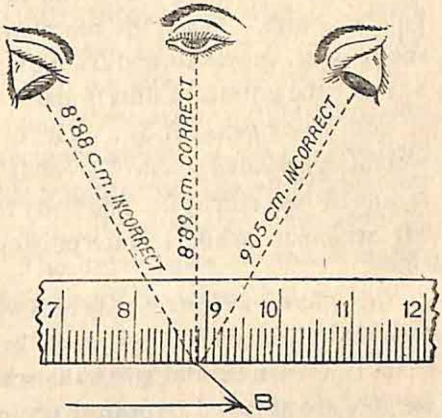

Fig. 3(c)

Owing to the thickness of the wood, the eye must always be placed vertically above the mark being read, otherwise *parallax error* may come in. Fig. 3(c) shows how oblique line of sight may give incorrect reading while vertical line of sight correct reading.

1·11. Measurement of mass :

For measurement of mass of a body, we generally employ an instrument known as common balance. With this instrument we determine the mass of a body in comparison with some standard weights The weighing scale widely used by shopkeepers in markets is a simplified form of the common balance. The following are the essential parts of a common balance [Fig. 4(a)] :—

(*a*) *Balance beam* (*AB*) *:* It is a long beam pivoted, at its centre, about an agate (a very hard semi-precious stone) knife-edge(C) which rests on another agate plate. The plate, again, is placed on the top a rod which is inserted in a vertical piller *H*. The rod can be raised up or lowered down by turning a knob *K*. When raised up, the balance beam will oscillate about the knife-edge *C* and when lowered down, the beam will come to rest. The knife-edge *C* is known as fulcrum.

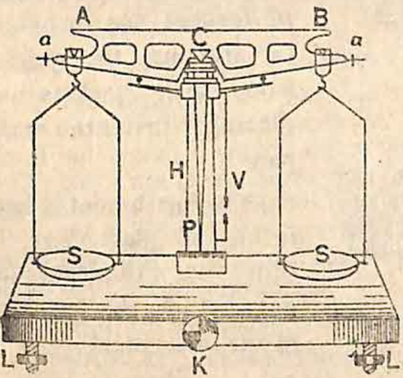

Fig. 4(a)

(*b*) *The Pointer* (*P*) :

A pointer with pointed end is fixed at the middle of the balance beam. When the beam swings freely, the pointed end of the pointer moves to and fro along a scale. When the beam is at rest, the pointer points 0-mark of the scale.

(*c*) *Scale pans* (*S*, *S*) : There are two scale-pans of equal weight suspended from the ends of the balance beam *AB* by means of two stirrups. The body to be weighed is placed on the left scale-pan while counterpoising weights are placed on the right.

(*d*) *Screw-weights* (*a*, *a*) : Two small screw-weights are provided at the ends *A* and *B* of the balance beam. If the balance beam is not horizontal when the scale-pans are empty, the screw-weights are adjusted in order to bring the beam in the horizontal position.

(*e*) *Plumb line* (*V*) : A plumb line is provided with all instruments in order to see whether the pillar *H* is vertical.

(*f*) *Weight box* : Although weight box is not a part of the balance proper, yet it is essential for the measurement of mass by a balance. Fig. 4(b). shows a weight box. Standard weights of different denominations, such as 100 grams, 50 grams, 100 milligrams, 20 milligrams etc. are arranged in different sockets in the weight box. A pair of forceps are provided for raising the weights from their sockets and placing them on the scale-pan.

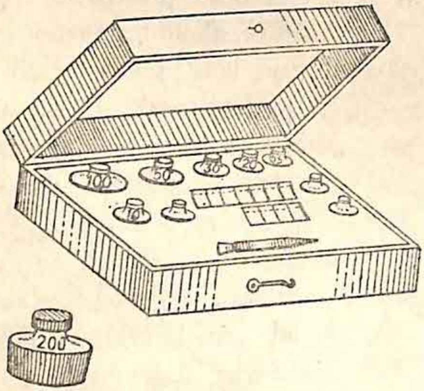

Fig.4 (b)

The instrument is kept inside a glass box in order that air current may not disturb the working of the instrument.

Ordinary use of the balance :

If the balance is accurate in all respects, then for ordinary measurement of the mass of a body, the following procedure is followed :

At first, the base of the balance is to be made horizontal by turning the levelling screws (L, L) at the bottom of the base. The plumb will, then, hang centrally through the ring. If, now, the knob K is slightly turned, the beam AB will be raised a little and the pointer will swing equally on either side of the zero mark. If it does not swing equally, the screw-weights (a, a) are to be adjusted. Then the substance to be weighed is placed on the left scale-pan and suitable weights from the weight box are put on the right scale-pan, by trial, till the pointer swings equally on either side of the 0-mark. The total mass of the counterpoising weights so placed is the mass of the substance.

Requisites of a good balance :

The following are the requisites of a good balance :

(i) The balance must be *sensitive i.e.*, the beam will get tilted if there is a slight difference in the masses of the two substances placed in the two scale pans.

(ii) The balance must be *true i.e.*, the balance beam must remain horizontal when the scale-pans are empty or when two exactly equal masses are kept on the scale pans.

(iii) The balance must be *stable i.e.*, the pointer, once set into oscillation, must come back to its position quickly.

(iv) The balance must be *rigid i.e.*, the parts of the balance must be strong enough to withstand the weight of the substance placed on the scale-pan.

Principle of weighing by a balance :

If the balance be otherwise true, then the standard weights placed on the right-hand scale-pan in order to make the balance beam horizontal when the unknown substance is on the left-hand scale pan, are taken as the mass of the substance. In Fig. 4(a), the mass of the unknown substance will try to rotate the beam about the fulcrum C in the counter clockwise direction while the counterpoising weights will do the same in the clockwise direction. Since the beam is horizontal.

Mass of the substance $\times AC =$ Counterpoising weights $\times BC$

As $BC = AC$, Mass of the substance $=$ Counterpoising weights.

1·12. Measurement of time :

Any event which repeats itself after a definite interval of time, may be used for the measurement of time. Before pendulum clocks were invented, many improvised arrangements like sand clock, water-clock, sun-dial, etc. were in use for the measurement of time. Fig. 5(a) shows a sand-clock used in ancient times. When sand runs out of the upper compartment into the lower one, a certain time is required. When the sand-clock is inverted, the running out of the sand from the upper to lower chamber takes place under the same condition, and hence in the same interval of time.

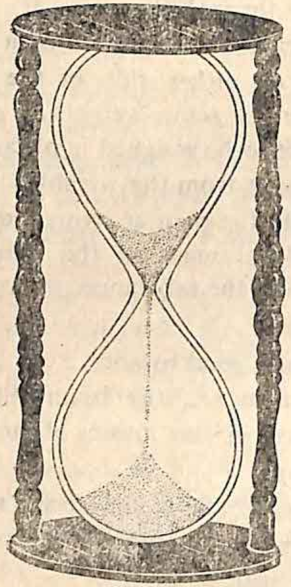

Fig. 5.(a): Sard Clock

At present, we use pendulum clocks, watches, etc. for the purpose of measuring time. Besides the usual clocks and watches there are chronometers, stop-watches, etc. in use. Stop-watches [Fig. 5(b)] have the advantage that they can be started or stopped at will. Stop-watches are sensitive enough to record $\frac{1}{5}$th of a second or even $\frac{1}{10}$th of a second. Stop-watches are generally used to record time-interval between events which happen in quick succession.

The principle of watch in measuring time is based on the use of an oscillating spiral spring furnished with a small fly-wheel called the balance wheel which is attached to an anchor type dead-beat escapement wheel.

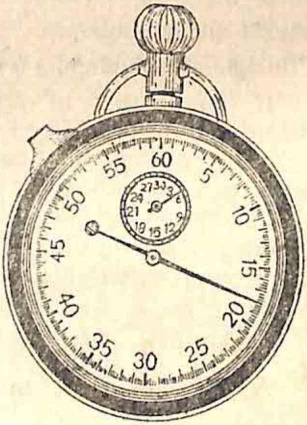

Fig. 5(b)

It is worthwhile to mention here that pendulum clocks came into use as a result of the discovery of simple pendulum by Galileo in 1581. A story is told of Galileo that when he was a boy of 17 years only, he was once attending a service in the Cathedral of Pisa and his attention was drawn to the swinging of a lamp which was suspended from the roof by a long chain. He counted his pulse-beats as a clock to determine how much time the lamp was taking for a complete swing. He was astonished to find that each complete swing—whether big or small—is taking equal number of pulse beats. Later he wrote in an article "I went so many times to the Cathedral, noticed the swinging of the lamp hundreds of times but never dreamt of learning that each swing takes place in a definite time."

Galileo was quick to realise the importance of this event and later on devised plans for the construction of a pendulum clock.

Exercises

1. What is a unit and what is its utility ? Explain the different systems of units.

2. Define the following quantities :—(a) Centimetre (b) Foot (c) Kilogram (d) Litre.

3. Express the following :—(a) a foot into centimetre (b) a metre into yard (c) a pound into kilogram (d) a second into day.

[Ans. (a) 30·48 (b) 1·09363 (c) 0·4536 (d) $\frac{1}{86400}$]

4. What are 'litre' and 'gallon' ? Express gallon into litre. [Ans. 4·54]

5. Measure the length of a table by a metre scale and by the palm of your hand. From this, determine how long your palm is, in metre. Now, directly measure the length of your palm by a metre scale and check your previous result.

6. The diametre of a fine wire is 0·008 mm. ; convert the diameter into metre. [Ans. ·000008]

7. What is the instrument for measuring the mass of a body ? Describe the instrument with a neat diagram. Explain the process adopted for measuring the mass.

8. What are the requisities of a good balance ?

9. What is a solar day ? What is its relation with solar second ?

10. Describe, in detail, the method of measuring the length of a substance by a scale. How will you avoid parallax error ?

11. What is a stop watch ? For what purpose is it used ?

12. Discuss, in brief, the principle of weighing a substance by a balance.

18. A few questions with possible answers are given below. Mark the answer with √ sign which you consider most accurate :

(i) What is the unit of volume in the C. G. S. system ? *Ans.* Cubic foot, cubic centimetre, gallon.

(ii) 5 c. c. of water is poured in a tea cup, in a glass and in a measuring cylinder. Will the volume of water remain same in each case ? *Ans.* No, yes, volume will be halved.

(iii) Can you call the unit of area a fundamental unit ? *Ans.* Yes, no, it is a derived unit.

(iv) A road is one mile long and another, 1 kilometre long. Are the roads equally long ? *Ans.* Yes ; First one is longer ; the second one is longer.

(v) What instrument will you use to determine the volume of a wooden block ? *Ans.* Balance ; Scale ; Measuring cylinder.

CHAPTER TWO

Matter and Energy

2·1. Introduction :

The earth of ours is full of objects. We see around us various kinds of objects. Table, chair, paper, pen, tree and such other objects which we see around us and which have got weight are called *bodies.* Some of them are solid, some liquid and some gaseous. Some, again, have good odour, some are odourless and some have bad odour. Many of them have colour, some may be colourless. How are these varieties of bodies formed ? How do they behave ? What are their structures ? These questions and numerous such other questions have tormented the mind of man from the very dawn of the civilisation. The inquisitive mind of man has, since then been trying to find answers to those questions. The informations gathered by them in course of time have given rise to different branches of science.

The material of which a body is made is called *matter.* For example, wood is a kind of matter but a chair made of wood is a body.

Besides matter, attention of man was also drawn to another important thing—viz., *Energy. Capacity for doing work is called energy.* The universe exists because energy exists. Without energy, the universe will come to a standstill. Our familiarity with different forms of energy grows through its association with material bodies. For example, heat is a form of energy. But heat has no shape, colour or odour like material bodies. To realise the existence of heat, we are to observe the physical changes which heat brings about in a body. Absorbing heat, a body changes its temperature, increases in length or in volume. In the same way, electricity is another form of energy. To realise its existence, we are to observe the effects when electricity flows through a body. When we touch a current-carrying wire or look at it, we do not feel the existence of electricity. But when electric current goes to a fan, the fan starts whirling and then we realise the existence of it. So, material bodies are essential for understanding energy.

Definition : Anything that is perceptible to our senses and has weight and free existence, is called a *body*. The material of which a body is made is called *matter*. Capacity of a body for doing work is called *energy*.

2·2. Different Physical States of Matter :

Some of the substances we meet are solids, such as iron, stone, wood, etc. Others are liquids such as oil or water. Still others are gases, such as air or steam. A solid substance can be considered as one that tries to keep a definite shape and a definite volume. A liquid has also a definite volume but no shape. Liquids will take up the shape of any container into which they are poured. A gas, on the other hand, has neither a definite shape nor a definite volume. Unlike water in a jar, a gas does not have a distinct surface.

Some common substances are mixtures of matter in several states. Fine sand or silt mixed with water will not settle out. It forms a *suspension*—a stable mixture of a solid and liquid. Ink is another example. Milk is an *emulsion*—globules of one liquid (fat) suspended in another (water). *Foam* is a gas suspended liquid.

Often, we know a single kind of matter in all three principal states. Water is a common example. Ordinarily, water is a liquid, but at low temperature, it goes into its solid state (called ice) and at higher temperatures, it becomes steam, which is the name for the gaseous state of water. We usually think of air as a gas, but at very low temperature, it turns into a bluish liquid. Iron commonly seen in the solid state, becomes a liquid in a foundry (at about 1700°C) and is a gas in the atmosphere of the sun where the temperature is many thousands of degrees.

Structure of Matter :

In whatever state a matter may exist, its ultimate constitution is the same. Matter is constituted of very tiny particles known as *Molecules*. Molecules can exist in free state and possess the property of the substance they constitute. Molecules are, again, made up of still smaller particles called *Atoms*. Atoms cannot exist in free state but can take part in a chemical reaction.

According to modern Physics, atom is not the smallest possible state of matter. It is constituted of very tiny particles, known as *electrons* which are charged with negative electricity. Electorns in an atom revolve round a positively charged nucleus which is made up of protons, neutrons, etc.

The structure of an atom may be likened with the structure of the solar system, where planets revolve round the sun in different orbits. The nucleus, in an atom, plays the role of the sun, while the electrons play the role of the planets. It is to be remembered that the nucleus is composed of positively charged *protons*, uncharged *neutrons* and some other fundamental particles.

2·3. Mass and weight of a body :

It has been mentioned before that mass of a body means the quantity of matter contained in it. The mass of an iron ball is the quantity of iron in it. Ordinarily, we do not make any difference between mass and weight of a body. We say, for example, that the mass of a body is 30 kilos and at the same time, we refer to the weight of the same body as 30 kilos. But they are entirely different quntities and their difference should always be borne in mind.

It goes without saying that all bodies have weight. We feel the weight when we place a body on our palm. Some bodies are so weighty that it is difficult to keep them on our palm or to lift them to a certain height. Do you know whence this weight of a body comes ? The earth attracts everybody towards its centre. *The weight of a body is the attractive force of the earth exerted upon the body*. More massive the body is, more is the attractive force on it and hence more is its weight. The weight of a bucket is more than that of a cup. If the bucket is full of water, its weight increases further.

Sir Isaac Newton, the celebrated Physicist, first noticed the attractive force of the earth. There is a very interesting story in this connection. In 1666, plague broke out in England in an epidemic form. Newton was then a student of Trinity College in Cambridge. All the students of the college, being afraid of the disease, fled away to distant places. Newton also came to his village home at Woolsthrope. One day he was reading a book below an apple tree in his garden. Suddenly an apple fell on the ground near him. Seeing the apple fall, he thought, "Why did the apple fall downwards ? It could have gone up ! Why do all substances, when let fall, come downwards ? Certainly the earth pulls everything towards it." He went on thinking about it and ultimately discovered the force of gravity. Just imagine a boy of 24 years, discovering a universal truth ! Newton later realised that gravitational attraction is applicable not only to substances on the earth but also to substances outside the earth.

The following simple experiment will show that the pull of the earth on different bodies is different *i.e.*, different bodies have different weights.

A spiral spring with a scale-pan and a pointer attached is held vertically by a clamp and stand against a scale graduated in grams. [Fig. 1]. The pointer moves along the scale when the spring elongates. Now, place a body on the scale-pan. The pointer will be found to slide down. Why does it do so ? The earth pulls the body downwards and the spring elongates in length. Hence, the pointer comes down. In other words, the spring elongates due to the weight of the body and hence the pointer slides down. If different bodies be kept on the scale-pan, the pointer will come down to different marks in the scale, showing that different bodies have different weights.

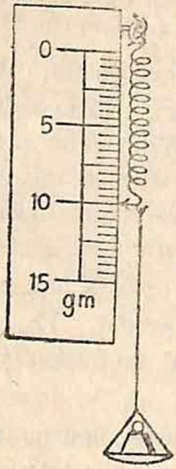

Fig. 1

There is an instrument known as *spring balance* (described later) which is used for determining the '*weight*' of a body. The instrument works on the basis of elongation of a spring due to weight.

From the above discussion, we may define weight of a body as the force with which it is attracted towards the earth. So, remember that *weight is essentially a force.*

Wherever a body is taken, its mass always remains the same. But the weight of a body changes when it is taken to different parts of the earth or to different altitudes. A body will gradually weigh less if it is taken from the south or the north pole to the equator. Again, the weight of a body diminishes when it is taken to the top of a mountain from the surface of the earth. Since mass of a body is not influenced by the motion, the position, the temperature, the electrical or the magnetic condition of the body, it is called an intrinsic property of the body. Weight, on the other hand, is not an intrinsic property of a body because weight is variable.

Spring balance : Fig. 2 shows a spring balance commonly used in a laboratory. It consists of a spring with a hook attached to the lower end, from which a weight can be suspended. A pointer is provided with the spring which moves over a graduated scale given on the casing, inside which the spring is fitted. When a weight is suspended from the hook, the spring gets elongated, elongation being proportional to the weight. So, the pointer attached to the spring moves over the graduated scale and by noting its position against the scale, we can directly measure the weight of a body.

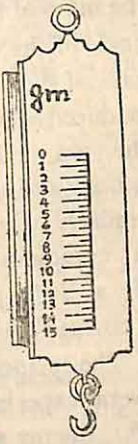

Fig. 2

2·4. Conservation of mass :

Scientists believe that *the total quantity of matter in the universe is fixed and it cannot be increased or decreased by human agency.* When we prepare a substance from one or more of other substances, either by chemical means or by any other means, the total amount of matter always remains the same.

Some day-to-day happenings may give rise to doubts in our mind about the truth of the principle. For example, when we burn a candle or a piece of paper, all that we are left with is a small amount of residue or ash. When we boil some water, it vanishes altogether. If we place a piece of camphor in open air, it will disappear after some time. As a matter of fact, we find wear and tear of all substances around us. Are these not destruction of matter ? Again, if we sow a small seed to-day, it will grow into a big tree years after. Burn a piece of magnesium wire in open air. You will get some white ash. If you weigh it, you will be astonished to see that its weight is greater than that of the magnesium wire ! Are these not creation of new matter ? Do these indicate a violation of the principle of conservation of matter ?

No, they do not indicate a violation of the principle. If we examine carefully we would see that when a candle burns, some gas is produced which gets mixed up with the surrounding air. The mass of the gas—if it could be measured—together with the mass of the residue left, will be found to be exactly equal to the mass of the candle. In the case of water also, some steam is produced whose mass is equal to the mass of the water taken. The seed, taking food from the soil and the air, gradually increases its mass and in course of time, is converted into a big tree. Similarly, when magnesium wire is burnt in air, oxygen from the air combines with magnesium which makes the mass of the ash left a little greater than the mass of the wire. So, we see that in no case mass is destroyed or created.

The principle has also been verified in the laboratory by several experiments.

2·5. Energy and its different forms :

It has been mentioned before that the capacity for doing work of an agent, is called its energy. In general, energy can be divided into seven forms.

(1) *Heat energy* : Heat can perform many works. For example, the heat of superheated steam can drive a steam-engine, the heat of the fuels like petrol, diesel, kerosene, etc. can work an automobile engine, an aeroplane engine or a pump. So, heat may be considered as a form of energy.

(2) *Light energy* : If we hold our hand near a glowing electric lamp, we feel hot. This means that light produces heat

energy. Hence, light is also a form of energy. That light is a form of energy is also proved by the fact that light can exert pressure.

(3) *Sound energy* : If a stretched string is plucked, it produces a sound. Here, the mechanical energy of the string is converted into sound energy.

(4) *Magnetic energy* : A magnet can perform some work by attracting a piece of soft iron. This proves that magnetism is also a form of energy.

(5) *Electric energy* : A charged body attracts an uncharged body or repels a similarly charged body. So, electricity may be reckoned as a form of energy.

(6) *Mechanical energy* : Energy manifested by mechanical process such as a carriage in motion, a shot fired from a gun, movement of muscles, etc. is known as mechanical energy.

(7) *Chemical energy* : When coal burns, it combines with the oxygen of the atmosphere, releasing chemical energy in the form of heat.

2·6. Transformation of energy :

The different forms of energy mentioned above are mutually interconvertible *i.e.*, one form of energy may be converted into another. As a matter of fact, all natural phenomena that we see around us, are manifestation of transformation of energy. Some simple illustrations of such transformation are given below :

Water flows from higher level to lower. At a higher level, the potential energy of the mass of water is converted into kinetic energy when it flows to lower level. This kinetic energy of flowing water, again, is utilised to generate electrical energy.

When electricity flows through the filament of an electric bulb, light is produced. This is an illustration of electric energy being converted into light energy.

In a steam-engine, steam is produced by heat, which keeps the locomotive running. Here, heat energy is converted into mechanical energy.

If we rub our palms together, heat is produced. Here, mechanical energy is converted into heat energy.

In this way, numerous cases may be cited illustrating transformation of energy of one form into another.

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2·7. Conservation of energy :

In all examples of transformation of energy, we find that a body loses energy while another gains an equal amount of it. As a matter of fact *we cannot create any new energy nor can we destroy any.* Scientists are of opinion that the total amount of energy in this universe is conserved. This is known as the principle of conservation of energy.

Exercises

1. Define : matter, body and energy. Explain the difference between matter and body.

2. In how many states can a matter exist ? Explain with illustration. What are atoms and molecules ?

3. What do you mean by mass and weight of a body ? What is their difference ? Why is mass called an intrinsic property of a body ?

4. Will it be correct to say that the mass of a body is 30 gms. and the weight of the body is 30 gms. ?

5. A body is taken from the surface of the earth to the top of a mountain. What change will you notice in respect of its mass and weight ?

6. Explain, with suitable illustrations, the principle of conservation of mass.

7. What is energy ? How many forms of energy are there ? Can one form of energy be converted into another ? Explain your answer with illustrations.

8. State the principle of conservation of energy.

9. Some incorrect statements are made below. Correct them wherever necessary :

(i) Energy manifests itself through a body. But when energy of one body is transformed into energy of another body, the first body always loses some energy.

(ii) Scientists are of opinion that the total amount of mass in the universe when it first came into existence, it still there. But observing the wear and tear of all substances around us, we may say that the mass of the universe is gradually diminishing.

(iii) When a particular body is weighed by a spring balance in different places, same result should be obtained, because the mass of the body remains unchanged.

CHAPTER THREE

Change of State

3·1. Introduction :

We know that matter can exist in three states *viz*, solid, liquid and gas. When a body changes from solid to liquid state or from liquid to gaseous or from gaseous to liquid state *i.e.*, from one state to another—it is called a change of state. For example, a piece of ice is the solid state of water. When it melts into water, it is converted into liquid state and when water passes into steam, the liquid state changes to gaseous state.

3·2. Melting and freezing :

Suppose, we have a piece of ice at −10°C. If the piece be heated, its temperature will rise. But when the temperature becomes 0°C, the piece of ice starts melting into water and the temperature is found to remain steady although heat is being supplied to it. This will continue till the whole of the mass of ice is melted into water. Afterwards, the temperature of water will increase.

Similarly, if a quantity of water is cooled, its temperature will fall until it attains a temperature of 0°C. After that, the temperature will remain steady at 0°C although heat is being extracted from the mass of water which will now start to solidify into ice. This state of affairs will continue till the whole of water is solidified into ice, after which, of course, the temperature of ice will fall.

That the above fact is seen only in the case of water, is not true. As a matter of fact, it is seen in all cases of substances which do not undergo chemical change due to heating. For example, wood or coal, which undergoes chemical change on heating, will not melt but will get charred.

As a general rule, therefore, we can say that the temperature of a solid body at first increases on being heated, until it attains a particular temperature when the body begins to melt. The temperature, then, remains steady till the whole mass of the body melts into the liquid state. This phenomenon is known as *melting.*

On the other hand, if heat be extracted from a quantity of liquid, its temperature will, at first, fall. Reaching a particular temperature, the liquid begins to freeze and the temperature remains steady till the whole mass is frozen into the solid state. This phenomenon is known as *freezing* or *solidification.*

3·3. Melting point and Freezing point of a substance :

Definition : *Melting point is the temperature at which a solid substance begins to melt under a given pressure and this temperature remains constant till the whole mass of the substance melts into the liquid state.*

Freezing point is the temperature at which a liquid begins to freeze under a given pressure and this temperature remains constant till the whole mass of the liquid freezes into the solid state.

Almost every substance has identical melting and freezing points. For example, ice melts at 0°C under normal atmospheric pressure while water freezes at the same temperature (0°C) under the same condition. But there are some substances like fat, wax, glass, butter—which pass through a vitreous state before melting. They do not exhibit any sharp melting or freezing point. Butter, for example, is found to melt somewhere between 28°C and 33°C and its freezing point lies somewhere between 23°C and 20°C. But it is to be always borne in mind that freezing or melting point of a substance is not a constant quantity ; it depends upon superincumbent pressure.

3·4. Effects of pressure on melting point :

It has been mentioned earlier that melting point of a substance depends upon pressure. The relation between the melting point and the pressure is as follows :—

(i) Substances like cast iron, ice, etc. *whose volume decreases on melting*, have their melting points lowered due to increase of pressure. Ice, for example, is found to have its melting point lowered by 0·007°C due to an increase of one atmospheric pressure. The reason is that increased pressure helps the contraction of volume more and therefore, lowers down the melting point.

(ii) Substances like wax, etc. *whose volume increases on melting*, have their melting point raised due to increase of pressure. In the case of wax, it has been found that its melting point is raised by 0·04°C due to an increase of one atmospheric pressure. Here also, the simple reason is that increased pressure causes hindrance to expansion of volume and therefore, raises the melting point.

3·5. Vapour and Vaporisation :

The aeriform state of a liquid is called its *vapour* and the process by which a liquid changes into vapour is called *vaporisation*.

It is relevant to mention here that *gas and vapour are not the same thing*. There is a difference between the two, although ordinarily we do not make any difference. Every student of science should bear in mind the difference between them.

Experiments show that vapour given out by a liquid cannot be liquefied again by simply exerting pressure, if the temperature of the vapour be kept at any value. For liquefaction the temperature must be of a particular value or below that value. This particular temperature is called the *critical temperature*, which is different for different gases. If the temperature of a vapour is below the critical value, it should be called a vapour, If, on the other hand, the temperature is above the critical value, it should be called a gas.

3·6. Different methods of vaporisation :

There are two ways in which a liquid can change into vapour, *viz*, (i) *evaporation* and (ii) *boiling or ebullition*.

(i) **Evaporation :** Evaporation is the process in which a liquid is slowly and gradually converted into corresponding vapour. Evaporation takes place at the surface of the liquid and under all pressures. Drying up of rivers and tanks in acute summer, disappearance of some water kept in an open vessel, drying up of water of wet clothes in sun, are examples of evaporation.

Rapidity or otherwise of evaporation depends upon the nature of the liquid and some other circumstantial factors. Pour equal quantities of water and spirit (methylated) in two shallow dishes and keep them in open place. Spirit will evaporate much earlier than water. This shows that the rate of evaporation depends upon the nature of the liquid. More volatile the liquid is, more rapid will be its rate of evaporation. Ether, alcohol, etc. are very volatile liquid.

You have certainly noticed that wet clothes dry up quicker in winter than in rainy season. The atmosphere, in rainy season, remains full of water-vapour which prevents rapid evaporation. Dryness of the atmosphere is, therefore, conducive to rapid evaporation.

While drinking hot tea or hot milk, some of you pour it in shallow dish because it helps to cool down the liquid. Do you know why ? When poured in a dish, the liquid spreads into a greater surface-area than when the liquid was in a cup. More the exposed surface-area, more quickly does it evaporate.

Wet clothes dry up quickly in a current of air ; sun-shine and not shadow helps to dry up wet clothes soon—these are our daily experiences, from which we know that air-current, higher temperature, etc. increase the rate of evaporation.

From the above examples, we can, therefore, summarise that following are the factors which influence the rate of evaporation of a liquid :—(*a*) *nature of the liquid*, (*b*) *presence of moisture in the atmosphere*, (*c*) *area of the exposed surface*, (*d*) *the air-current*, (*e*) *the temperature of the liquid.*

3·7. Boiling or ebullition of a liquid :

The process in which a liquid very rapidly changes into corresponding vapour is called boiling or ebullition. Boiling takes place from all points throughout the liquid but at a particular temperature, depending upon the surrounding pressure.

The constant temperature at which a liquid boils under normal atmospheric pressure is called its *normal boiling point.*

Experiment : Take some water in a flask whose mouth can be closed by a rubber stopper. Through one hole in the stopper insert a thermometer (*T*) and through the other a small bent tube (*E*). See that the bulb of the thermometer remains a little above the water in the flask [Fig. 1]. Fix up the flask in a suitable stand as shown in the figure and place a piece of wire-gauge below the flask. Now, heat the flask slowly by means of a burner.

When the water is heated a little, vapour will be found to rise from the surface and the dissolved air will form small bubbles which will come out of water and will be deposited on the wall of the vessel. The thermometer will show a gradual rise of temperature. When the mercury touches 70°C/80°C mark of the thermometer, bubbles of water-vapour will be formed at the bottom of the flask. These bubbles will, then, rise upward and coming in contact with comparatively colder water, will collapse. At this stage, a simmering sound will be heard. Finally when the thermometer shows a temperature of 98°/99°C, these bubbles rise to the top and burst at the surface, causing a turmoil in the whole mass of water. At this time, profuse steam will come out of the bent tube *E* and the thermometer will record a steady temperature. When this happens, we say that the liquid is boiling. The temperature will remain steady as long as the boiling continues.

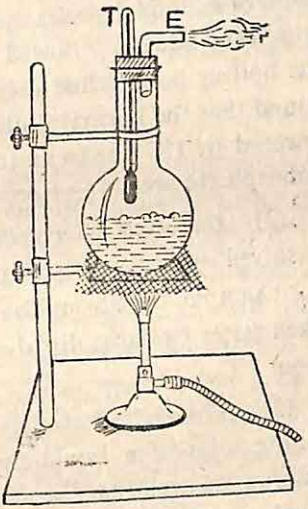

Fig. 1

3·8. Difference between boiling and evaporation :

The following are the points of difference between boiling and evaporation :—

(i) Boiling is a very rapid process while evaporation is a slow one.

(ii) Boiling takes place throughout the mass of the liquid but evaporation takes place from the surface only.

(iii) Boiling takes place at a particular temperature, depending upon the superincumbent pressure but evaporation takes place at all temperatures and under all pressures.

3·9. Factors influencing the boiling point of a liquid :

The following factors influence the boiling point of any liquid :—

(1) *The superincumbent pressure of the liquid* : The boiling point of a liquid depends upon the superincumbent pressure under which the liquid is allowed to boil. Increase of pressure raises the boiling point while decrease of pressure lowers it. It has been found that the normal boiling point of water (100°C) is raised or lowered by 1°C due to an increase or decrease of every 27 m.m. of atmospheric pressure respectively.

(2) *Dissolved impurities in the liquid* : Liquids with dissolved impurities have higher boiling point than pure liquids. For example, the normal boiling point of pure water is 100°C but when water contains dissolved salts, its boiling point is raised by about 9°C.

(3) *Material of the vessel* : It has been found that the boiling point of a liquid depends, to a certain extent, upon the material as well as the cleanliness of the vessel in which the liquid is boiled. For example, water boils at a higher temperature in a glass vessel than in a copper vessel. The boiling point is still further raised if the glass vessel is neat and clean.

3·10. Condensation :

The process of converting a vapour into its liquid state is called *condensation*. It has been mentioned before that vapour is formed when a liquid is heated. So, if the vapour be condensed into liquid again, the vapour is to be cooled. Such condensation of water-vapour in the atmosphere into water drops is frequently seen in nature.

During autumn if you look at the blades of grass in the morning you will find thousands of glittering dew-drops on them. The water-vapour present in the atmosphere is cooled during night and is deposited as dew-drops on grass and leaves.

Clouds are formed as a result of condensation of water-vapour in a large mass of air and the water particles present in the cloud, when cooled further, appear as rain drops.

Turn the lid of the vessel in which rice is being cooked. You will find copious water-drops condensed on it.

Drop a piece of ice in some water taken in a glass vessel. You will find that in a short time, water-drops are deposited like fog on the outer wall of the vessel. The vessel is cooled due to the presence of ice and the vessel, in turn, cools the water-vapour present in the air surrounding the vessel. The water-vapour, therefore, condenses on the outer wall of the vessel.

3·11. Latent heat :

When a body is heated, its temperature rises. With the help of a thermometer, we can watch the rise of temperature and can realise that the body is absorbing heat. But if a piece of ice at 0°C, be likewise heated, the thermometer will, at first, show no change of temperature although ice, absorbing heat, will melt gradually. This state of affairs will continue till the mass of ice melts completely. After this, of course, the temperature of water obtained by the melting of ice, will increase. What happens then, to the quantity of heat supplied to the ice, during its melting ? It helped ice to melt but has no external manifestation. In this way, whenever a solid melts into liquid, it absorbs an amount of heat which a thermometer cannot record. This heat remains hidden or latent in the body and hence, is known as **Latent heat.**

In the same way, if a certain quantity of water is heated, its temperature steadily rises until the temperature reaches 100°C. The rise of temperature can easily be recorded by a thermometer. At this stage, the temperature will remain steady at 100°C, although water absorbing heat, starts to boil. In other words, this heat does not increase the temperature of water but helps it to change from liquid state to vapour state. In this way, any liquid during its evaporation, will absorb some heat which a thermometer connot record. It is, also, called latent heat.

In the first case, the latent heat is known as the *latent heat of fusion* while that in the second case, *the latent heat of vaporisation.*

In general, we can say that a body during its change of state will either absorb or give out a quantity of heat which has no external manifestation. This heat is called latent heat.

The latent heat of fusion of ice is 80 calories. This means that 1 gm. of ice at 0°C requires 80 calories of heat for conversion into water at 0°C.

Likewise, the latent heat of vaporisation of water is 537 calories, which means that 1 gm. of water at 100°C will be converted into steam at 100°C by absorbing 537 calories of heat.

3·12. Cold caused by evaporation :

In order to change into vapour, a liquid requires some latent heat. If this heat be not supplied from any external sources, the liquid will extract it from its own body or from its surroundings, causing a fall of temperature *i.e.*, cooling of its own body or of its surroundings. Such cooling effect due to evaporation is frequently met with.

(a) If a little methylated spirit is spilt on the hand, it evaporates rapidly and the hand feels very cold. To change from liquid to vapour, the spirit requires latent heat. This it obtains from the hand which thus loses heat and cools. For the same reason, in high fever, an *eau-de-cologne* bandage or simply water-bandage is pasted on the forehead of the patient. Eau-de-cologne or water, on evaporation, takes up latent heat from the body and consequently the temperature of the patient comes down.

(b) Perspiration is the body's own method of maintaining a constant temperature. When perspiring heavily after exercise or hard work, it is unwise to stand about in a draught, because overcooling may result from evaporation. The resulting chill may lower the body-resistance to infection.

Dogs, who do not perspire from the skin, hang out their tongues during hot weather in order to obtain a cooling effect.

(c) In hot weather, drinking water is kept in an earthenware vessel. Water oozes out through the pores of the vessel and gets evaporated. The latent heat necessary for the evaporation is supplied by the vessel. The resultant cooling keeps the temperature of water inside the vessel several degrees below the outside temperature. Metal or glass vessel cannot keep water so cool because those vessels have no pores and therefore, water cannot ooze out and evaporate. Old earthenware vessel also cannot keep water cool as efficiently as the new one, because its pores become choked with dust and other particles.

(d) It is unwise to sit in draught with wet clothes on. Water from the clothes will quickly dry up, taking latent heat from the body. The resultant cooling will lower the body-resistance and the person may catch cold.

Exercises

1. What do you mean by melting and freezing ? What do you understand by the statement 'the melting point of platinum is 1755°C' ? Is melting point of a substance identical with its freezing point*?

2. What is the effect of pressure on melting point ? Explain with examples.

3. What are evaporation and boiling ? What are their differences ?

4. What are the factors which influence the rate of evaporation of a liquid ? What is the difference between a gas and a vapour ?

5. What is boiling point ? What is its relation with the superincumbent pressure ?

6. *Answer the following questions :—*

(i) Why does fanning cause comfort in hot weather ?

(ii) An earthenware vessel keeps water cool but a metallic vessel does not. Why ?

(iii) Why do people use '*khaus-khaus*' in doors and windows in summer ?

(iv) Why does a person catch cold if he sits in draught with wet clothes on ?

(v) Why does an *eau-de-cologne* bandage on forehead cause comfort in high fever ?

7. What is latent heat ? Is this heat detectable by a thermometer ? What is the value of the latent heat of fusion of ice ?

8. Cite some examples of condensation of water-vapour present in atmosphere.

9. What is latent heat of vaporisation ?

10. Some questions are given below. Write 'Y' against those whose answers are 'Yes' and 'N' against those whose answers are 'No'.

(i) Although temperature remains steady as long as the melting or freezing of a substance continues, is it correct to say that the substance does not absorb or give out any heat during that time ? —

(ii) Does every substance have definite melting or freezing point ? —

(iii) Is the melting point raised due to an increase of pressure for those substances whose volume expands on melting ? —

(iv) Can the vapour issued from the surface of water be called a gas ? —

(v) Will a liquid be quickly evaporated if the air-current over it is minimised ? —

(vi) Will the boiling point of a liquid be lowered if the superincumbent pressure is lowered ? —

PHYSICS

CHAPTER ONE

Rest and Motion

1·1. Introduction :

In our every-day life we see various bodies in motion as well as in rest. Buildings, trees, roads, towers, statues, etc., are bodies at rest while running trains, horses, cars, etc., are bodies in motion. When you walk to your school from your home, you are in motion. A man cycling along a road is in motion. But a parked car is at rest. Numerous such examples of rest and motion are known to you. If you watch a moving body, you will see that *it is changing its position with respect to time.* But a body at rest, will not change its position. Your school building will be in the same position to-morrow or for all the time to come, where you find it to-day. For this reason, a moving body means one which changes its position with respect to time while a body at rest means one which does not change its position with respect to time.

But you know that the earth rotates round its own axis once in 24 hours and round the sun once in 365 days [Fig. 1]. How can buildings, trees, etc. situated on the moving earth,

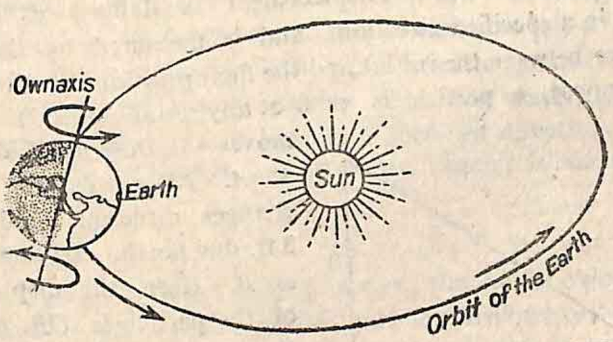

Fig. 1

remain at rest ? When the astronauts move far away from the earth in space-craft, they see all the buildings, trees, etc.

moving speedly along with the earth. As a matter of fact, nothing in this universe, is at rest. *Absolute rest* in something unknown to us.

If this be so, why do we say that buildings, trees, towers, etc. are at rest ? What sort of rest it is ? Ordinarily, if a body does not change its position relative to its surrounding objects, we say the body to be at rest. On the other hand, if a body changes its position relative to its surrounding objects, we say the body to be in motion, These are, therefore, *relative rest* and *motion.* So, when we say that a body is at rest or in motion, we always mean the relative rest or relative motion in respect of the surrounding objects, assuming that the earth is at rest. In this sense, trees, buildings, etc. are at rest, because they never change their positions relative to the surrounding objects.

1·2. Translation and some terms in connection with translation :

When a body moves in a straight line, its motion is called *translation.* Examples of translation are not far to seek. It a stone be dropped from a certain height, it falls along a straight vertical line. So, the motion of a falling stone is a case of translation. If a car moves along a straight road, its motion may be called a translatory motion. The following are a few important terms in connection with translatory motion :

(*a*) **Displacement** ; Displacement is defined as distance moved in a specified direction, and is measured by the linear distance between the initial and the final positions of the body.

Suppose, a particle is at *O* at any instant (Fig. 2). It then moves 4 ft. due east and arrives at *A*. After this, the particle changes direction and moves 3 ft. due north, arriving finally at *B*. Here, the displacement of the particle is *OB*, although it went along the path *OAB*, beacuse *O* is the initial and *B* is the final positions of the particle. The magnitude of the displacement $= OB = \sqrt{(4)^2+(3)^2} = 5$ ft,

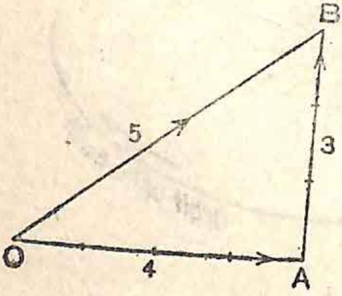

Fig. 2

Displacement, therefore, is a quantity which carries a sense of direction as well as of magnitude. Such quantities, which have direction as well as magnitude are called *vector quantities.*

(*b*) **Speed** : Speed is defined as the rate of change of position with time, *i.e.*, the distance travelled in 1 second. So, to find speed, we require the total distance travelled and the total time taken for it, and in that case,

$$\text{Speed}=\frac{\text{Distance travelled}}{\text{Time taken}}=\frac{D}{t}.$$

For example, if a man travels a distance of 1 mile in a cycle instead of walking, he will take shorter time. Again if he travels the distance in a car, he will take still shorter time for the journey. We say that the speed of cycling is greater than the speed of walking. On the other hand, the speed of a car is greater than the speed of a cycle.

It is to be noted that nowhere above have we mentioned the direction of motion. As a matter of fact, speed has magnitude and no direction. Such quantities which have only magnitude and no direction, are called *scalar quantities.*

If the body travels equal distances in equal intervals of time, its speed is said to be *uniform.* If it does not do so, the speed is *variable* or *non-uniform.*

(*c*) **Velocity** : Velocity is defined as the rate of change of position in a specified direction *i.e.*, the distance travelled in a specified direction in 1 second. Velocity is a vector quantity.

Velocity may be *uniform* or *variable.* If a body travels equal distances in equal intervals of time in a specified direction, its velocity is uniform. If it does not do so, the velocity is variable or non-uniform.

Units of velocity :

C. G. S. System : In the C. G. S. system, the unit of velocity is 'centimetre per sec.' *i.e.*, if a particle travels a distance of 1 cm. in 1 sec. in a specified direction, its velocity is taken as unit velocity in the C. G. S. system.

F, P. S. System : In the F. P. S. system, the unit of velocity is 'foot per sec'. *i.e.*, if a particle travels a distance of 1 foot in

1 sec., in a specified direction, its velocity is taken as unit velocity in the F.P.S. system.

Examples : (1) If a train travels a distance of 10 miles in 10 minutes, what is its speed ?

Ans. Here, the total distance travelled $= 10$ miles

$$= 10 \times 1760 \times 3 \text{ ft.}$$

and ... ... time taken $= 10$ mints. $= 10 \times 60$ sec.

$$\therefore \text{ speed} = \frac{\text{Total distance}}{\text{Total time}} = \frac{10 \times 1760 \times 3}{10 \times 60} = 88 \text{ ft/sec.}$$

(2) A motor car travels due south for 1 hour and covers a distance of 15 kilometres. What was its velocity ?

Ans. Here, the total distance travelled due south

$$= 15 \text{ kilometres} = 15 \times 1000 \times 100 \text{ cms.}$$

and ... ... time taken $= 1$ hour $= 60 \times 60$ sec.

$$\therefore \text{ Velocity} = \frac{\text{Distance travelled}}{\text{Time taken}} = \frac{15 \times 1000 \times 100}{60 \times 60}$$

$$= 416{\cdot}6 \text{ cm./sec.}$$

Difference between speed and velocity :

Velocity is a vector quantity, whereas speed is a scalar quantity. In ordinary conversation, we use the word 'velocity' very often in place of speed. In science, however, it is important to distinguish between these two terms. The distinction will be well understood if we take an illustration.

For example, if a car were travelling at a steady speed of 20 miles per hour along a perfectly straight road, it would be correct to say that it had a velocity of 20 miles/hr. due south, or whatever the direction of the road might be. On the other hand, if the car were travelling round a bend with constant speed, its direction of motion would be continuously changing. Hence it would be incorrect to say, in this case, that the velocity of the car was 20 miles per hour.

(*d*) **Acceleration :** When a train leaves a station, it moves with gradually increasing velocity and finally it attains a constant velocity. At the initial stage, when the velocity of the train was changing, the train may be said to have an acceleration. When a particle moves with gradually increasing velocity, *its rate of change of velocity with time is defined as its acceleration.* Let us take an illustration.

Suppose, at any instant, a particle has a velocity of 32 ft./sec. It moves with accelerating motion and after 10 seconds, the velocity changes to 52 ft./sec. and after further 10 seconds, the velocity again changes to 72 ft./sec. The particle is moving in this manner with gradually increasing velocity. Here, it is seen that after each 10-second interval the particle is changing its velocity by 20 ft./sec. Hence, the rate of change of its velocity, in other words, its acceleration$=\frac{20}{10}=2$ ft. per sec. per sec.

Note that the word *'per sec'* is appearing twice in the unit of acceleration—first to denote the velocity and the second to denote the rate of change of velocity.

In the C.G.S. system, the unit of acceleration is cm./sec./sec. and in the F.P.S. system ft. per sec. per sec.

It is to be remembered that like velocity, acceleration may be *uniform* and *variable.*

(*e*) **Retardation or Deceleration :** When a particle moves with gradually decreasing velocity, its rate of change of velocity with time is defined as its retardation. Retardation may also be called negative acceleration. If brakes be applied to a moving train, its velocity gradually decreases and finally it comes to rest. In this condition, the train may be said to have a retardation. As before, let us take an illustration.

Suppose, at any instant, a particle has a velocity of 32 ft./sec. After 2 seconds, its velocity is found to be 28 ft./sec. and a further period of 2 sec. after, its velocity decreases to 24 ft./sec. The particle is moving in this manner with gradually decreasing velocity. Here, it is seen that after each 2-second interval, the particle is changing its velocity by 4 ft./sec. Hence, the rate of change of its velocity, in other words, its retardation$=\frac{4}{2}$ $=2$ ft,/sec./sec.

The units of acceleration and retardation are identical.

Examples : A particle, starting from rest, attains a velocity 40 ft./sec. in 4 sec. What is its acceleration ?

Ans. The change of velocity of the particle=40 ft./sec. and the time reqd.=4 sec.

$$\text{Hence, its acceleration}=\frac{\text{Change of Velocity}}{\text{Time}}=\frac{40}{4}$$

$$=10 \text{ ft./sec./sec.}$$

1·3. Newton's laws of motion :

In 1687 Sir Isaac Newton published a book named "*The Principia*" in which he discussed the motion of a body—particularly that of the planets and other heavenly bodies. In the first part of the book, Newton summed up the basic principles of motion in three laws, known as the laws of motion.

First law : *Everybody continues in its state of rest or of uniform motion in a straight line unless compelled by some external impressed force to act otherwise.*

Second law : *The rate of change of momentum of a body is proportional to the applied force and takes place in the direction in which the force acts.*

Third law : *To every action, there is an equal and opposite reaction.*

Let us now discuss the laws one by one.

1·4. Discussion of first law :

From a discussion of first law of motion, we come to know of the following :—(i) *inertia of matter* and (ii) *definition of force.*

Inertia of matter : The first law states that if a body be at rest, it will continue to be at rest or if a body be in motion, it will, likewise, continue to be in motion in a straight line. The tendency of material bodies to remain at rest or to continue any motion they are performing is referred to as the inertia of the material bodies.

Inertia may be divided into two parts, viz., (i) *inertia of rest* and (ii) *inertia of motion.*

It is not difficult to have a conception of inertia of rest because it is a matter of common experience that objects, at rest do not begin to move of their own accord. If we place an object at a certain place we expect it to remain there unless a force is applied to it. So, inertia of rest is a common-sense matter.

But it is not immediately obvious that a body moving with uniform velocity in a straight line tends to go on moving for ever without coming to rest. A ball rolled along the ground moves for some time but finally comes to rest. The ball does not move for ever. How is, then, the truth of inertia of motion defended ? It is very difficult to defend because no one has yet found a means of eliminating the various outside forces which can retard a moving body. The ball does not move for ever because external forces like friction, air-resistance, etc., retard the motion of the ball. We know if we roll a ball on a smooth floor or an ice, it will go further than it goes on the ground, because a smooth floor or ice offers less friction than ground. It would, therefore, be reasonable to suppose that if air-resistance, friction, etc., could be eliminated, the ball would go on moving in a straight line for ever. In this way, we can have a conception of the inertia of motion.

Illustrations of inertia of rest :

(*a*) If a car, carrying passengers, suddenly starts moving, the passengers have a tendency to fall back. This is an example of inertia of rest. When the car is at rest, the whole body of a passenger is at rest. With the sudden start of the car, the lower part of the body, which is in contact with the car, moves forward but the upper part, due to inertia of rest, continues to be at rest. The passenger, consequently, leans backwards.

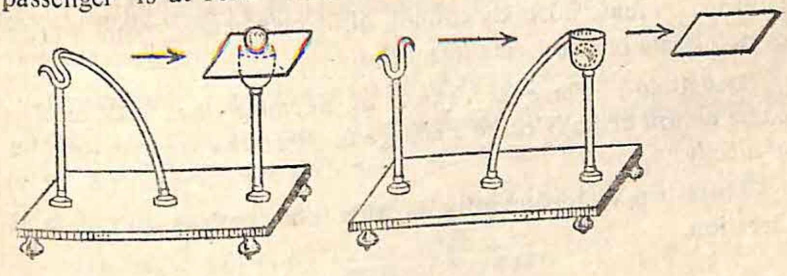

Fig. 3

(*c*) A cup is placed on a vertical stand [Fig. 3]. A small ball rests on a card placed over the cup. If, now, a spring is pulled aside and let go, the spring will flick the card away. The ball will be found to drop neatly into the cup. The card moves away so quickly that the ball cannot partake of the motion of the card. The ball continues to be at rest and drops into the cup.

Illustrations of inertia of motion :

(*a*) A person riding a bicycle along a level road does not come to rest immediately he stops pedalling. The bicycle continues to move forward due to inertia of motion but eventually it comes to rest due to retarding forces of air-resistance and friction.

(*b*) When a train or motor car is suddenly stopped, the passengers are shot forward because their bodies tend to maintain their forward motion.

Similarly, if a passenger alights from a moving car absent-mindedly, he is thrown in the forward direction. When inside the moving car, his whole body is in motion. But as soon as his feet touch the ground, the lower part of his body comes to rest but the upper part, due to inertia of motion, continues its forward motion.

1·5. Force :

Form the first law, we also know that an external force must be applied on a body in order to change its state of rest or of motion. A material body cannot, of its own accord, come to rest or accelerate or decelerate its motion.

Definition : *Force is defined as anything that will tend to cause or will actually cause a change in the state of rest or motion of a body.*

Force is a vector quantity, because it has both magnitude and direction.

1·6. Discussion of second law :

From a discussion of second law of motion, we come to know of measurement of force and of a relation between the force and acceleration or retardation. Before we take up the subject, we shall discuss something of "*momentum*" of a body.

Definition : The momentum of a body is defined as the product of its mass and its velocity and is a measure of its quantity of motion.

If 'm' be the mass of a body and 'v' its velocity, then its momentum$=m\times v$. For example. the momentum of a motor car weighing 2000℔ and moving with a velocity of 44 ft./sec., is $2000\times 44=88{,}000$ units.

Now, if a force be applied on a body moving with uniform velocity, the velocity of the body, depending upon the direction of the force, may increase or decrease. In any case, there will be a change of momentum of the body. The second law states that the change of momentum will depend upon the applied force. The change of momentum will be greater if the force increases and vice versa.

Besides, the change of momentum takes place in the direction in which the force acts. From these considerations, we may write, *Applied force=mass of the body×acceleration of the body.*

This gives us a measure of force.

1·7. Units of force in different systems :

Dyne is the unit of force in the C. G. S. system. It is the force which produces an acceleration of 1 cm./sec^2, when it acts on a mass of 1 gm.

Poundal is the unit of force in the F. P. S. system. It is the force which produces an acceleration of 1 ft./sec^2, when it acts on a mass of 1 ℔.

Remember, 1 poundal$=13{,}800$ dynes (nearly),

Example : What will be the acceleration produced when a force of 500 dynes acts on a mass of 175 gms. ?

Ans. We know, force=mass of the body×acc^n. of the body

$$\therefore \text{ acc}^n\text{. of the body}=\frac{\text{force}}{\text{mass of the body}}=\tfrac{500}{175}$$

$$=2{\cdot}86 \text{ cm./sec}^2$$

1·8. Discussion of third law :

Suppose, we have two bodies A and B. If the body A exerts a force on the body B, then according to the third law, the body

B will, in return, exert an equal and opposite force on *A*. If the force of *A* on *B* is called *action*, then the force of *B* on *A* may be called *reaction*. If a weight of 5 kilos rests on a table (Fig. 4), there will be an upward reaction R=W=5 kg. from the table.

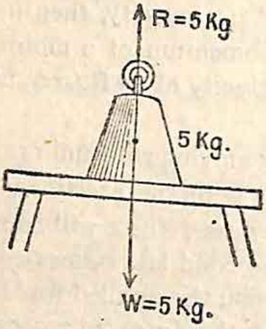

Fig- 4

A simple experiment may be done to show that action and reaction are equal and opposite. Fix the hook of a spring-balance with that of another and stretch the two balances by hand equally in the opposite directions. It goes without saying that the readings of the two balances will be same. Let it be 5 ℔s. [Fig. 5(up)]. Now, fix one spring balance against a rigid support, say a wall, and stretch the other equally as before, with hand [Fig. 5(down)]. In this case, also, the two balances will show 5 ℔ reading as if the spring-balance, fixed to the wall, is being stretched by someone equally in the opposite direction, as before. Here, this spring-balance is experiencing reaction, and the two readings being same, we can say that action and reaction are equal and opposite.

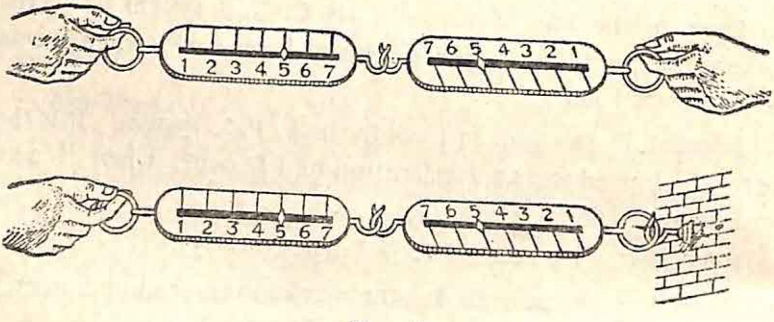

Fig- 5

It is to be remembered that *action and reaction act upon two different bodies*. So, the question of establishment of equilibrium does not arise although the two forces are equal and opposite. Equilibrium is established, it may be recalled, when two equal and opposite forces act on the *same* body.

Illustrative examples of third law are not far to seek. We discuss some of them below :—

(*a*) When a man jumps ashore from a boat, the boat is pushed back. The person exerts a force on the boat, which, therefore, is pushed back and the boat exerts an equal and opposite reaction on the person, which takes him to the shore.

(*b*) When a bullet is fired from a gun, the gun-man feels a push backward. It is the reaction of the bullet on the gun.

(*c*) The rockets which form a familiar feature of fire-work displays contains chemicals which burn to produce high-velocity blast of hot gas, which issues out through a nozzle in the downward direction. The tremendous reaction it produces hurls the rocket skyward.

Exercises

1. Define and write the units of the following quantities :—
   (i) speed, (ii) velocity, (iii) acceleration, (iv) retardation.

2. Write brief answers to the following questions :—

(*a*) What is the distinction between speed and velocity ?

(*b*) Why do you mention 'time' twice in the units of acceleration ?

(*c*) Why is retardation called negative acceleration ?

3. A particle moves with a velocity of 5 ft./sec. for 5 minutes. How far does it go ? [Ans. 1500 ft.]

4. The speed of a particle changed from 1 ft./sec. to 1 mile/hr. in 1 minute. Find the acceleration of the particle, taking yard as the unit of length and minute as the unit of time. [Ans. $9\frac{1}{3}$ yds./min^2]

5. Two trains are moving with same velocity in the same direction along two parallel rail lines. How will a passenger in one train appear to move relative to another passenger, sitting front to front, in the other train ?

6. State Newton's laws of motion.

7. Explain how from the first law, we get a definition of force and from the second law, a measure of force.

8. What do you mean by inertia of matter ? Explain with illustrations. Why do persons lean backward while alighting from a running car ?

9 Define force and explain its units in different systems.

10. A body of mass 50 gms, acted on by a force, moves with an acceleration of 5 cm./sec^2. What was the magnitude of the force ? [Ans. 250 dynes]

11. What are action and reaction ? Do they act on the same body ? How would you prove that action and reaction are equal and opposite ?

12. Some statements are made below. Write "C" against those which you consider correct and "W" against those which you consider wrong :—

(a) Having both magnitude and direction, retardation is a vector quantity. —

(b) Trees, buildings, etc., are at rest and this is absolute rest. —

(c) The velocity of a body rotating along a circle is an example of uniform velocity. —

(d) The product of mass and velocity of a body gives us momentum of the body. —

(e) Action and reaction can produce equilibrium condition in a body. —

13. Fill the blanks in the following sentences :

(a) Everybody — in its state of rest or of — motion in a — line, unless compelled by some — impressed — to act otherwise.

(b) The rate of change of — of a body is — to the applied — and takes place in the — in which the — acts.

(c) — is defined as anything that will tend to cause or will actually cause a — in the state of — or of — of a body.

CHAPTER TWO

Work, Power and Energy

2·1. Work :

In our daily life, we come across numerous examples of work. When labourers carry loads, horses or bullocks pull carriages, boatman rows a boat, engine pulls a train, they all do some work. As a matter of fact, all of us do some work always. Even when we are asleep, our heart does work by constant beating. In ordinary conversation, the word "*work*" refers to almost any kind of physical or mental activity but in science it is used in a special sense which will be clear from the following illustration.

Suppose some masons are carrying bricks up a ladder for the construction of a building. If we wish to compare the work done by one mason with that done by another, we might think it sufficient to count the bricks each has carried. A little thought will show that this may not give a true comparison. If one mason has carried 100 bricks to a height of 40 ft. and the other has also carried 100 bricks but only to a height of 20 ft., the former has done more work than the latter.

Evidently there are two factors which must be considered when estimating an amount of work done by the masons. When climbing with a load of bricks, a mason would be exerting a certain force to overcome the weight of the bricks and his own weight. The magnitude of this force has certainly something to do with the work done, but the distance through which the bricks are raised must also be taken into account.

As a matter of fact, in estimating an amount of work done, the product of the force applied and the displacement of the point of application of the force is to be determined.

Definition : Work is said to be done when the point of application of a force moves and is measured by the product of the force and the distance moved in the direction of force. In other words,

Work done = Force × displacement of the point of application of the force in the direction of the force.

From the above definition, it is clear that whatever may be the magnitude of the force, unless the point of application of the force moves, no work will be done. For example, you may perspire profusely in your attempt to move a huge block of stone but no work is considered to have been done, unless the stone moves. The great mythological Serpent *Basuki* who supports the world on its hood, must have become very tired by now, but scientifically speaking, it did no work at all, since it merely exerted an upward force on the world and did not move it.

Now, if the point of application of the force moves *in the direction of the force*, work is said to be done by the force. For example, if a body be dropped from a certain height, the body falls downwards due to the attractive force of the earth. Here, the body moves in the direction in which the attractive force acts. So, in this case, work is said to be done by the attractive force of the earth.

If, on the other hand, the point of application of the force moves *opposite to the direction of the force*, work is said to be done against the force. Thus, if a heavy load is lifted up, work is said to be done against the attractive force of the earth.

2·2. Different units of work :

(i) **Absolute units** : In the C. G. S. system, the absolute unit of work is *erg*. If the point of application of a force of 1 dyne moves through a distance of 1 cm. in the direction of the force, the work done is 1 erg.

In the F. P. S. system, the absolute unit of work is *foot-poundal*. If the point of application of a force of 1 poundal moves through a distance of 1 ft. in the direction of the force, the work done is 1 foot-poundal.

(ii) **Gravitational units** : In the C. G. S. system, the gravitational unit is *gram-centimetre*. It is the work done in raising a mass of 1 gm. vertically through 1 cm. against the attraction (*i.e.*, gravity) of the earth.

Remember, 1 gram-centimetre=981 ergs.

In the F. P. S. system, the gravitational unit is *foot pound.* It is the work done in raising the mass of 1 ℔ vertically through 1 ft. against the attraction of the earth.

Remember, 1 foot-pound=32 foot-poundals.

(iii) **Practical units :** The absolute unit 'erg' in the C.G.S. system is widely used. But in estimating work in some cases, erg is found to be a very small unit. In those cases, a bigger unit—the practical unit—is used. It is known as *Joule*.

Remember, 1 Joule$=10^7$ ergs.

Further, we know, 1 foot-poundal = 1 poundal × 1 foot.

Now, 1 poundal=13800 dynes (approx) and 1 ft.=30·48 cm.,

$\therefore$ 1 foot-poundal$=13800\times30{\cdot}48$ ergs
$=4{\cdot}2\times10^5$ ergs (nearly),

Also, 1 foot-pound=32 foot-poundals
$=32\times4{\cdot}2\times10^5$ ergs$=1{\cdot}35\times10^7$ ergs (nearly)
$=1{\cdot}35$ Joules (,,)

2·3. Power : Power is defined as the rate of doing work.

Suppose, two gardeners are raising water from a well with the help of a bucket. It is evident that more powerful gardener will raise greater number of bucketful of water in a certain time than the other. In other words, more powerful person is capable of doing more work in a certain time. So, to find power, we are to determine the ratio of the total work done and the time taken for it.

$$\text{Power }(P)=\frac{\text{Work done }(W)}{\text{Time taken }(t)}.$$

2·4. Different units of power :

(i) **Absolute units ;** In the C. G. S. system, the absolute unit of power is *erg/sec*. If an agent does 1 erg of work in 1 sec., its power is 1 erg/sec.

In the F. P. S. system, the absolute unit of power is *foot-*

poundal/sec. If an agent does 1 foot-poundal of work in 1 sec, its power is 1 ft.-poundal/sec.

(ii) **Practical units :** In the C. G. S. system, the practical unit of power is *Watt.* If an agent does 1 joule of work in 1 sec. its power is 1 watt.

$\therefore$ 1 watt = 1 joule/sec. = 10^7 ergs/sec.

In some cases, another bigger unit, known as *Kilowatt* (Kw) is used. 1 Kw. = 1000 watts.

Generally, watt is used to denote the power of electrical apparatus and mechineries.

In the F. P. S. system, the practical unit of power is *Horse-Power.* It denotes 550 ft.-lbs. of work in 1 second.

$\therefore$ 1 H. P. = 550 ft.-lbs/sec.

Remember, 1 H. P. = 746 joules/sec. = 746 watts (nearly).

2·5. Energy :

Anything which is able to do work is said to possess energy. Ordinarily, we say that a man with greater energy is capable of doing more work.

Definition : Energy is defined as the capacity to perform work. The measure of energy of a body is the amount of work performed by it. Its unit is same as the unit of work.

It has been mentioned earlier that energy can be divided, in general, into seven different forms. We are, at present, concerned with mechanical energy which is, again, divided into two parts, viz., (i) *Kinetic energy* and (ii) *Potential energy.*

2·6. Kinetic energy :

On the bed of a swiftly running stream, sand and stones may often be observed being rolled along by the water. Evidently the water must possess energy to enable it to do the work of moving these bodies and, as it would not move them if it were still, it possesses this energy because of its motion.

The work of driving the machinery of a wind mill is performed by air hitting the sails. Here, too, the air possesses energy because of its motion.

When a bullet is fired from a gun, the bullet can penetrate glass *i.e.*, it can do some work. This work the bullet can perform because it has considerable motion.

So, we see that motion endows a body with some energy which is known as kinetic energy.

Definition : Kinetic energy is the energy which a body has by reason of its motion.

It may be shown that the kinetic energy of a moving body is given by, K. E$=\frac{1}{2}\times$mass of the body$\times$(velocity of the body$)^2$.

2·7. Potential energy :

You must have seen a toy motor-car. When it is wound up, it moves. There is a spring in the toy and it possesses energy when it is in a condition of being wound up. When the spring uncoils, it performs a certain amount of work in driving the motor-car through a certain distance. So, it is clear that the spring possesses energy when it is given a configuration other than its standard configuration.

When a hammer is allowed to fall from a certain height on a peg, the peg penetrates into the ground. But the peg will not penetrate if the hummer is just placed on the peg. This shows that the hammer acquires the ability to do some work only when it is raised to a certain height.

Definition : Potential energy is the energy which a body possesses by virtue of its position or configuration.

Different types of machine

2·8. Introduction :

When work becomes hard, man applies his intelligence to devise means for making the work easy. As a result, we find different forms of machines. The age in which we live is frequently called the machine age and there are few forms of work to-day in which machinery is not used.

Many people think of a machine as being a more or less complicated piece of mechanism which includes gear, wheels, levers, screws and so on. But, however complex a machine may be, its various parts can always be shown to be applications of a limited number of basic mechanical principles.

In general, *a machine may be described as any arrangement which enables a force* (known as *effort*) *applied to it at one point to overcome a force* (known as *resistance*) *at some other point.* Machines are usually designed so that a small effort may overcome a large resistance or load. The advantage that we get from a machine is called *mechanical advantage.* If the load or resistance overcome be W and the effort applied be P, then

$$\text{mechanical advantage} = \frac{\text{load}}{\text{effort}} = \frac{W}{P}.$$

The following are known as simple machines—(i) *Levers,* (ii) *Inclined plane* and (iii) *Wheel and axle.*

2·9. Levers :

A lever consists of a stout rod capable of turning about a fixed point, which is called the *fulcrum.* Effort is applied at one end of the rod which overcomes the load applied at some other point. Levers may be classified into three classes.

(*a*) *Lever of the first class* : Fig. 1 shows the action and application of this lever. AB is the lever rod. C is the fulcrum which is situated not at the middle of the rod but nearer one end. In fig. 1, the fulcrum is nearer B-end. The effort (P) is applied at the end A, while the load (W) at the end B. Since AC is greater than BC, W will be greater than P. This means that a big load may be raised by a small effort. So, the lever of the first class has sufficient mechanical advantage. Crowbar (Fig. 1), balance beam

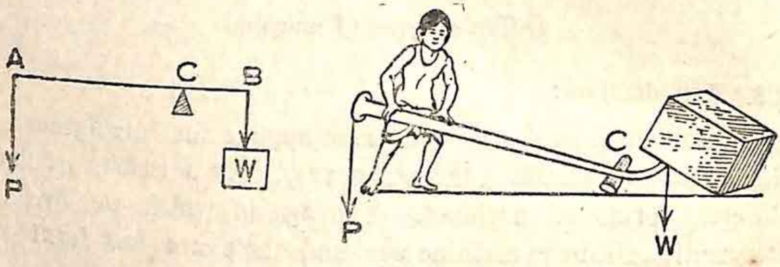

Fig. 1

Railwaymen's lever, are examples of lever of the first class. A pair of scissors may be cited as a good example of double lever of the first class.

(b) *Lever of the second class* : Fig. 2 shows the action and application of this lever. In this lever, the fulcrum C is situated at one end of the rod and the effort P at the other. The load W is applied in between them but nearer to the fulcrum C.

Here, since the arm AC is always greater than the arm, BC, the load will be always greater than the effort. In other words, in the lever of the second class, a small effort will always overcome a heavy load. The lever has, therefore, mechanical advantage. A wheel-barrow [Fig. 2], an oar of a boat, cork-squeezer, etc., are examples of the lever of the second class. A nut-cracker is an example of double lever of this class.

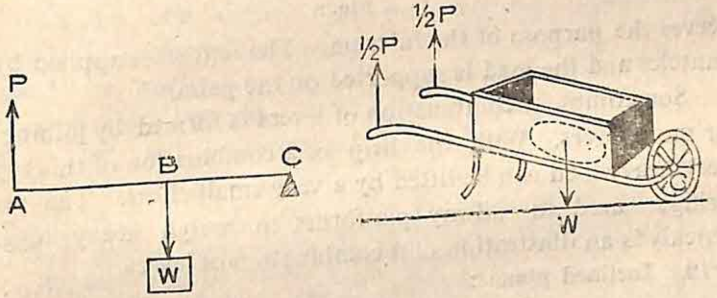

Fig. 2

(c) *Lever of the third class :* Fig. 3 illustrates the action of this lever. Here, also, the fulcrum C is situated at one end of the rod but the load W is applied at the other end. The effort P is applied somewhere between them but nearer to the load. Here, since AC is always smaller than BC, the load will be always smaller than the effort. In other words, a small load is overcome by a heavy effort. For this reason, *this lever has no mechanical advantage*.

Sometimes it may so happen that the load cannot be lifted directly by applying force nor is it convenient to use the levers of the first two classes. In such cases, the lever of the third class is used although it has no mechanical advantage.

Human arm [Fig. 3] is an example of the lever of the third class. The bone of the arm is the lever rod and the elbow-joint serves the purpose of the fulcrum. The effort is applied by the muscles and the load is supported on the palm.

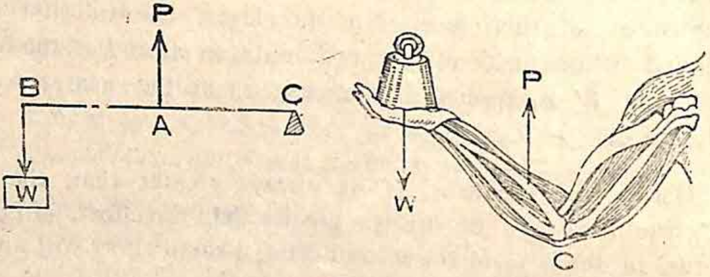

Fig. 3

Sometimes, a combination of levers is formed by joining two or more levers. With the help of a combination of this type, a very heavy load can be lifted by a very small effort. The 'weigh bridge' used in railway platforms to weigh heavy luggages quickly is an illustration of a combination of levers.

2·10. Inclined plane :

Those of you who have gone to a hill-station must have noticed that it is easier to negotiate a peak along a sloping way than along a steep way. For this reason, the vehicles going to a hill-station move along a spiralling way avoiding steep road as far as practicable. It is our common experience that we feel difficulty to ascend to third or fourth floor of a building if the staircase is steep.

From all these illustrations, it is clear that a heavy load may be raised more easily by pulling it along a sloping surface than by lifting it vertically. Workmen loading barrels on to a lorry usually roll them up a sloping plank [Fig. 4].

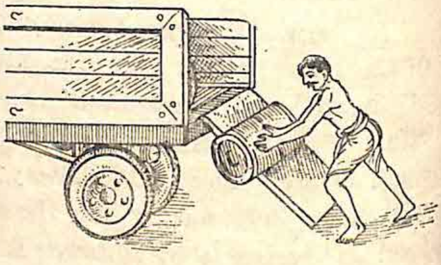

Fig. 4

Any smooth plane, like a wooden plank, kept at a certain angle with the horizontal may be called an inclined plane and since it gives mechanical advantage, it may be regarded as a machine.

Let AB be an inclined plane, inclined to the horizontal AC at an angle BAC [Fig. 5]. If, now, a load W be hauled up to B by applying an effort P along the plane, then the work done=force× displacement$=P\times AB$.

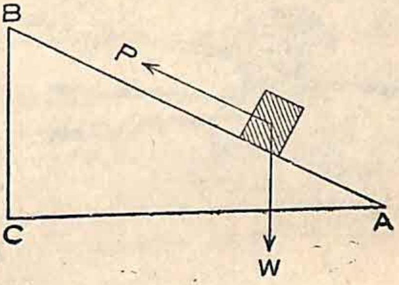

Fig. 5

As a result of this work, we may say that the load has been raised vertically through a height CB and the work done for that purpose=weight ×height$=W\times BC$. Since these two are equal,

$$P\times AB=W\times BC.$$

From Fig. 5 it is apparent that AB is much greater than BC. Hence P is much less than W *i.e.* by applying small effort, big load may be raised.

It is believed that large blocks of stone used in the construction of the Sun-temple of Konarak, Jagannath temple of Puri, and Egyptian pyramids were raised into their position by dragging them, on rollers, up a long ramp of earth, which was removed when the construction was over.

2·11. Wheel and Axle :

Fig. 6 shows the arrangement of the wheel and axle. It comprises of two cylinders of different diameters rigidly fixed on the same axle. The cylinder of larger diameter (B) is known as '*the wheel*' and the cylinder of smaller diameter (A) as '*the axle.* They have common axis about which they can rotate simultaneously. Two strings are wound over the two cylinders in opposite directions. Load(W)

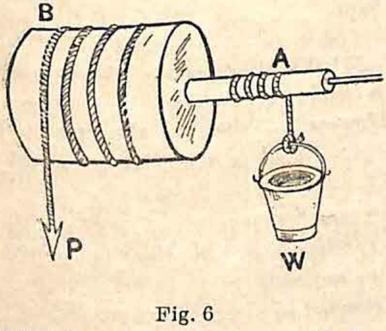

Fig. 6

is applied at the free end of the string wound over the axle while effort (P) at the free end of the string attached with the wheel. When the string of the wheel uncoils, that of the axle coils up and vice-versa.

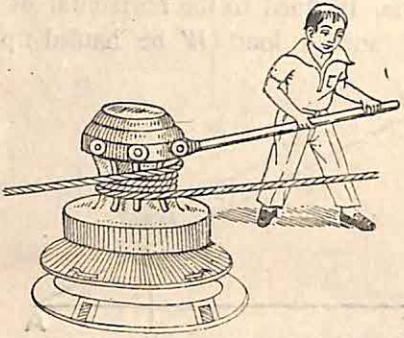

Fig. 7

In this case, because the diameter of the wheel is greater than that of the axle, the machine gives mechanical advantage.

The capstan used on board a ship is a good example of the wheel and axle [Fig. 7]. Here the length of the rod attached to the capstan gives the radius of the wheel while the radius of the barrel equals that of the axle.

One point must be remembered in connection with a machine. Using a machine, we make work easier *i.e.*, we over come a heavy load by applying a small effort but we do not gain anything in respect of energy. The output of energy from a machine is same as the input of energy ; the machine cannot create any extra energy. As a matter of fact, the output of energy is rather less than the input because some energy is dissipated to overcome the friction between the different parts of the machine.

Exercises

1. What do you mean by 'work' ? What is the difference between 'work done by a force' and 'work done against a force' ?

2. What do you mean by 'work' and 'power' ? What are their practical units in the C. G. S. and F. P. S. systems ? What is the relation between the units in the two systems ?

3. Explain the different units of 'work' and 'power'.

4. A man is keeping himself stationary with respect to the shore swimming against a strong current. Is he doing any work ?

5. What is 'horse-power' ? What is it equal to ?

6. What is energy ? Explain, with suitable illustrations, the difference between the two types of mechanical energy.

7. What is the advantage of a machine? Is there any mechanical advantage in the lever of the third class? Under what circumstances is it advantageous to use this lever?

8. How many classes of levers are there? Mention the characteristics and practical application of each class of lever.

9. What class of machine does a capstan belong to? Give a description of it.

10. A few questions are given below. Write 'Y' against those whose answers you consider 'yes' and 'N' against those whose answers you consider 'no' :—

(i) Is work equal to the product of the force and the displacement of its point of application? —

(ii) Are ordinary meaning and the scientific meaning of the word 'work' identical?—

(iii) A carriage is running with a constant speed along a level road and no net force is acting on the carriage. Is any work done on the carriage in this condition?

(iv) Is every class of lever equally advantageous?—

(v) Is the energy of an oscillating pendulum a mixture of kinetic and potential energy?—

11. A few questions with their probable answers written against them are given below. Mark the answer with √ sign that you consider most reasonable :—

(i) Where is the fulcrum situated in a lever of the first class?

Ans. At one end of the rod, somewhere at the middle of the rod.

(ii) What is an illustration of a lever of the second class?

Ans. Crowbar, an oar, human arm.

(iii) What is that quantity whose unit is horse-power?

Ans. Energy, work, power.

(iv) How many ergs is 1 joule equal to?

Ans. 10^5 ergs, 10^7 ergs, 32 ergs.

(v) When a body is raised to a certain height, what kind of energy does it possess?

Ans. Kinetic energy, potential energy.

(vi) What is capacity for doing work called?

Ans. Power, force, energy.

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CHAPTER THREE

Heat

3·1. Heat :

All of us have some idea of heat. We all know that we get heat if we put on a fire, that summer is hotter than winter, that we get heat when sun rises in the morning, etc. Since we cannot give any shape or size or colour to a bit of heat like a solid or liquid, we are to realise heat with the help of a material body. Our common experience says that when a body absorbs heat, it becomes hot and when it gives up heat, it becomes cold. In a simple way, therefore, heat may be looked upon as an external agency whose absorption turns a body hot and extraction turns a body cold.

3·2. Sources of heat :

The main source of heat on the earth is the sun. Solar heat is stored up in various indirect ways, which gives us electrical energy, chemical energy, etc. as other sources of heat.

The sun radiates enormous amount of heat energy in space and we on the surface of the earth, receive a very small part of it. Yet, this small part of solar heat that we receive is the ultimate source of all terrestrial energy. Living beings as well as plants exist because the earth receives heat from the sun.

Direct application of solar heat for the welfare of human beings, in an extensive manner, has not yet been made possible. But in a small way, the solar energy has been utilised in keeping the building hot, for irrigational works, in driving small machines, in running electric batteries, etc. You may have heard of 'Solar Cooker' which can cook food taking heat from the solar radiation.

3·3. Temperature :

All of us have a feeling of hotness and coldness. We feel cold if we touch a piece of ice while we feel hot if we touch a glowing electric bulb. Temperature, in a simple way, may be said to be the feeling of hotness or coldness of a body. The temperature of a body which feels hot to touch is said to be higher than that of a body which feels cold to touch.

At the same time, it should not be confused with the quantity of heat in a body. The temperature of a burning match-stick is much higher than the temperature of a bucketful of boiling water, but the water contains a much larger quantity of heat than the match-stick.

The term '*temperature*' is so important in the subject 'Heat' that it requires a detailed study.

Suppose, a hot iron ball is put inside a bucketful of cold water. It will be found that the iron ball is gradually becoming cooler and the water warmer. It is never seen that the hot ball gets hotter and the cold water colder. The reason is that the temperature of the hot ball, initially being higher than the temperature of water in the bucket, the ball gives up some heat which the water in the bucket absorbs.

Definition : *The temperature of a body is its degree of hotness and expresses such a thermal condition of the body that determines whether the body will give up heat to or absorb heat from another body.*

In this respect, temperature may be compared with the level of water. We know, water flows from upper level to lower level. Reverse flow never takes place. In other words, the level determines the direction of flow of water. Similarly, temperature determines the direction of flow of heat.

When a body *A* hands over heat to another body *B*, we say that the temperature of *A* is higher than that of *B*. If the heat flows from *B* to *A* we say that the temperature of *B* is higher than that of *A*.

Difference between heat and temperature :

(1) Heat is a form of energy whereas temperature represents a thermal condition of a body.

(2) When a body absorbs heat, its temperature rises and when it gives up heat, its temperature falls. Heat is, therefore, the *cause*, the temperature its *effect*. .

(3) The difference that exists between a quantity of water and its level is the same as the difference between a quantity of heat and temperature.

(4) Equal temperature of two bodies does not signify equal quantities of heat in them. In the same way, equal quantities of heat in the two bodies, do not signify equal temperature of the two.

3·4. Thermometer and thermometric scales :

Rough ideas of differences in temperature may be obtained from our sense of touch, but it is evident that accurate or correct comparisons when temperature differences are small cannot be made in that way. For accurate and correct measurement of temperature, instruments are necessary. Such instruments are known as thermometers.

The thermometer which uses mercury as its liquid, is called a *mercury thermometer*. Such type of thermometers are widely used. All of you must have seen a mercury thermometer. It consists of a small bulb [Fig. 1] at the end of a capillary tube of uniform bore. The bulb and part of the tube are filled with mercury and the upper end of the tube is sealed so that mercury will neither spill nor evaporate from the tube. In order that mercury may expand freely into the upper part of the tube, air has been removed from it so that it is a vacuum. A scale of degrees is marked on the stem to indicate temperatures. If the bulb of the thermometer is kept in contact with the body whose temperature is to be measured, the mercury will expand and the mark to which the level of mercury reaches gives the temperature of the body.

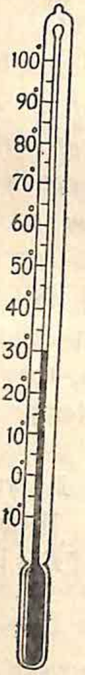


Fig. 1

Temperature Scale :

To obtain a scale of temperature, the melting point of ice is taken as the *lower fixed point* and the boiling point of water as the *upper fixed point*. Two thermometric scales are in common use in our country, *viz.* (i) **Celcius** or **Centigrade** scale and (ii) **Fahrenheit** scale.

(i) *Celcius or Centigrade scale* : According to this scale, the melting point of ice is taken as 0° degree and the boiling point of water as 100° degree. The space between these two points is divided into 100 equal divisions and each division is called one centigrade degree. [Fig. 2]. This scale was first devised by Celcius, a Sweedish astronomer.

(ii) *Fahrenheit scale* : According to this scale, the melting point of ice is taken as 32° degree and the boiling point of water as 212° degree. The space between these two points is divided into 180 equal divisions and each division is called one Fahrenheit degree. This scale was first devised by a British Scientist named Fahrenheit.

If C and F represent a particular temperature in Centigrade and Fahrenheit scales respectively, then their relation is,

$$\frac{C}{100}=\frac{F-32}{180} \text{ or } \frac{C}{5}=\frac{F-32}{9}.$$

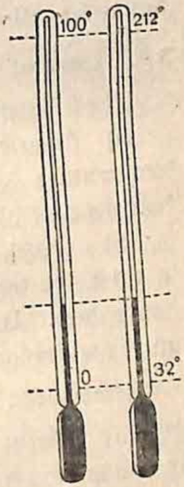


Fig. 2

It is worthwhile to mention here that the clinical thermometers used by the physicians are graduated according to the Fahrenheit scale. In this thermometer there are graduations from 95° to 110° Fahrenheit because the temperature of living human beings lies within this range. Clinical thermometers, graduated in centigrade scales, have recently been introduced.

3·5. Quantities of heat :

Many of the measurements made in Physics are made from direct reading instruments. Thus, mass is measured by a balance, temperature by a thermometer, length by a metre scale and so on. But there is no simple instrument which enables us to measure a quantity of heat directly. For measurement of heat we are to observe the effects it produces in a material body. For example, when a body is heated, its temperature changes.

Again it is seen that rise or fall of temperature of a body due to a certain quantity of heat given to or abstracted from it, depends upon the material of the body. Thus, if H be the quantity of heat given to a body and if the consequent rise of temperature be 't', then,

$H \propto t$ when the mass (m) of the body remains constant
$H \propto m$,, ,, temperature (t) ,, ,, ,,
i.e., $H \propto m.t.$ when both mass and temperature vary.
or, $H = S.m.t.$ where S is a constant of proportionality. This constant is known as **specific heat**. For this reason, to measure a quantity of heat, we require the mass of the body, the rise or fall of its temperature and a new quantity *viz.*, the specific heat. But before discussing it in detail, we must define units of heat.

3·6. Units of measurement of heat :

The following are the different units of measurement of heat.

(i) *Calorie* : It is defined as the heat required to raise the temperature of 1 gm. of water through one degree centigrade. It is the unit of heat in the C. G. S. system.

(ii) *British thermal unit* : It is defined as the heat required to raise the temperature of 1℔ of water through one degree Fahrenheit. It is the unit of heat in the F. P. S. system. This unit is widely used by engineers in England.

Remember : 1 B. Th. U.=252 calories.

(iii) *Therm* : It is the commercial heat unit used mainly in connection with the gas supply in England.

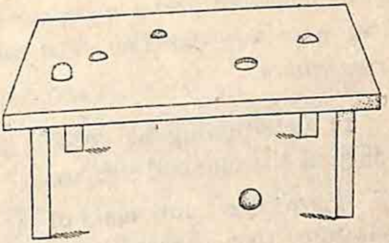
Remember : 1 therm=100,000 B. Th. U.

It may, therefore, be said that 1 therm is the heat required, to raise the temperature of 100,000 ℔s. of water through one degree Fahrenheit.

3·7. Specific heat :

If we take equal quantities of different materials like lead, iron, copper, etc. and heat them to have equal rise of temperature, we will see that different materials are taking different quantities of heat. So, it is clear that the capacity of absorbing heat by different materials does not depend only upon the mass or temperature. The following experiments will make the idea clear.

Experiments : (1) Take several balls of equal mass but of different materials like lead, copper, iron, etc. Raise the temperature of the balls equally by applying heat and put them quickly on a plate of wax. One will drop down by melting the whole thickness of the plate while others may melt half the thickness and so on (Fig. 3). It shows that although the balls have the same mass and same fall of temperature (because each one had same initial temperature and arrived at the same final temperature corresponding to the melting point of wax), yet, they gave up different quantities of heat. So, the amount of heat given up by a body does not depend only upon the mass or temperature change in the body.



(2) Take equal masses of water and milk in two identical kettles and warm them by the same burner. It will be found that the rise of temperature of milk is always greater than that of water. It, therefore, shows that although milk and water have the same mass and have the same supply of heat, yet their rise of temperature is not the same. So, the rise or fall of temperature of a body does not depend only upon the mass of the body or the quantity of heat supplied to it.

From the above experiments we can conclude that the quantity of heat absorbed or given out by a body does not depend only upon the mass or change of temperature of the body but also upon another property of the substance known as its *specific heat*.

In the first experiment, the different balls gave different quantities of heat because lead, copper, iron, etc., have different specific heats. In the second experiment, on the other hand, milk and water had different rise of temperature because they absorbed different quantities of heat due to difference in their specific heats.

Definition of Specific heat :

It is the ratio of the quantity of heat necessary for a given mass of a substance for a given rise of temperature to the quantity of heat necessary for the same mass of water for the same rise of temperature.

In determining the specific heat of a solid or liquid, water is taken as the standard substance.

If we take unit mass of the substance and unit rise of temperature, then, according to the above definition, we can write,

$$\text{Sp. heat} = \frac{\text{Heat necessary for unit mass of the substance for unit rise of temp.}}{\text{,, ,, ,, ,, ,, ,, water ,, ,, ,, ,, ,,}}$$

So, sp. heat is a ratio of two heats, and therefore, it is a pure number, having no unit.

In the C. G. S. system, 1 gm. represents unit mass and 1°C unit rise of temperature. So, in this system,

$$\text{Sp. heat} = \frac{\text{Heat necessary for 1 gm. of a substance for 1°C rise of temperature}}{\text{,, ,, ,, ,, ,, water ,, ,, ,, ,,}}$$

But, the denominator of the above ratio denotes 1 calorie. So, *the specific heat of a substance means the amount of heat in calories necessary for 1 gm. of the substance for 1°C rise of temperature*. Thus, the specific heat of copper is ·09 means that 1 gm. of copper will require ·09 calorie of heat for 1°C rise of temperature.

Likewise *the sp. heat of a substance also means the amount of heat, in B. Th. U., necessary for* 1℔ *of the substance for* 1°F *rise of temperature.* Thus the specific heat of copper is ·09 also means that 1 ℔ of copper will require ·09 B. Th. U. of heat for 1°F rise of temperature.

3·8. Amount of heat either absorbed or given out by a body for a rise or fall of temperature :

If S be the specific heat of a substance, then from the definition of specific heat, we can write,

1 gm. of the substance for 1°C rise or fall of temperature, will absorb or give out S calories of heat
$\therefore m$ ... ... mS ... ...
or, ... t°C ... $m.S.t.$ ... ...

If H be the amount of heat, then $H=m.S.t$, calories.

i.e., the amount of heat absorbed or given out
=mass of the body × its sp. heat × its rise or fall of temperature.

If, the initial temperature of the body be t_1°C and the final temperature, after absorbing heat, be t_2°C, then the rise of temperature=(t_2-t_1)°C and in that case,

$$H=mS(t_2-t_1) \text{ cal.}$$

In the same way, if the initial temperature of the body be t_1°C and the final temperature after giving out heat, be t_2°C, then, the fall of temperature=(t_1-t_2)°C and in that case,

$$H=m.S.(t_1-t_2) \text{ cal.}$$

Examples : (1) A body made of copper weighs 180 gms. The sp. heat of copper is ·09. How much heat is necessary to raise the temperature of the body from 25°C to 95°C ?

Ans. Here, $m=180$ gms ; $S=·09$; $t_1=25$°C ; $t_2=95$°C.

We know, $H=m.S.(t_2-t_1)$
$=180\times ·09\times(95-25)=180\times ·09\times 70=1134$ cal.

(2) How much heat will be necessary to raise the temperature of 2·5 ℔s. of alcohol from 68°F to its boiling point 173°F ? Sp. heat of alcohol=0·6.

Ans. $m=2·5$ ℔s ; $S=0·6$; $t_1=68$°F ; $t_2=173$°F.

We have, $H=m.S.(t_2-t_1)$
$=2·5\times 0·6\times(173-68)$
$=2·5\times 0·6\times 105=157·5$ B. Th. U.

3·9. Heat is a form of energy :

It is our common experience that friction produces heat. Different parts of a big machinery are lubricated by suitable lubricating oil in order to avoid friction and consequent production of heat. You have read stories of how ancient people used to kindle fire by rubbing two pieces of stone. Many such examples may be given where heat is produced by mechanical energy like friction.

If pressure be applied on an enclosed gas, some work is done which produces heat. While pumping air in a bicycle tyre or a football bladder you must have no iced such heating.

When electric current flows through the filament of an electric bulb, light and heat are produced. From this, we understand that electrical energy can be transformed into light or heat energy.

Examples of transformation of kinetic energy into heat are frequently met with. When a man runs swiftly, his body becomes warm. When blacksmith hammers a piece of iron, the piece becomes hot. These are examples of transformation of kinetic energy into heat.

From the examples mentioned above, it may be said that *heat is a form of energy* because it is obtained as a result of transformation of other forms of energy.

3·10. Relationship of heat with work :

When any form of energy is transformed into another, no energy is lost according to the principle of conservation of energy. If heat be a form of energy and if heat be obtained by transformation of mechanical energy, it must obey the above principle. It may, therefore, be said that the mechanical energy spent and the heat energy obtained must be proportional to each other. In other words, *to obtain a definite quantity of heat, a definite amount of work is to be done.* In 1847, James Prescott Joule, the celebrated scientist first determined the above relation and verified it experimentally too.

If W amount of work is transformed into H quantity of heat, then, according to the above relation, $W \propto H$ or, $\frac{W}{H} = \text{a constant.}$ This constant is called *mechanical equivalent of heat* and is generally, expressed by the letter 'J'.

$$\text{So, } \frac{W}{H} = J,$$

$$\text{or, } W = J.H.$$

Remember : In the C. G. S. system, $J = 4{\cdot}2 \times 10^7$ ergs/cal. This means that $4{\cdot}2 \times 10^7$ ergs of work are necessary to produce 1 calorie of heat.

Exercises

1. What is heat ? Name the different sources of heat. 'Main source of terrestrial heat energy is the sun'—explain the statement fully.

2. What is the difference between heat and temperature ? Which instrument is used to measure temperature ?

3. How many thermometric scales are there ? Explain them. In which scale is the clinical thermometer generally graduated ?

4. In trying to cleanse a clinical thermometer, a nurse, through mistake, put it in boiling water. The thermometer, afterwards, was found to be damaged. Why ?

5. Define the following terms :—(a) Calorie (b) British thermal unit (c) Therm (d) Sp. heat.

6. 'The specific heat of lead is 0·03' – what dose this statement mean ?

7. Two identical kettles containing equal masses of water and milk are put side by side, on the same fire. Rise of temperature was found to be quicker in milk than in water. Why ?

8. What factors are necessary in order to measure a quantity of heat ?

9. A pound of iron and a pound of lead, both at 100°C, are kept on a block of ice. Iron melts more ice than lead. Why ?

10. How much heat is necessary to warm 50 gms. of water from 15°C to 40°C ? [Ans. 1250 cal]

11. 1 lbs. of a particular type of fuel gives, on burning, 10,000 B. Th. U. of heat. How much fuel will be necessary to raise the temperature of 50 gallons of water through 100°F ? 1 gallon of water weighs 10 lbs. [Ans. 5 lbs]

12. 'Heat is a form of energy'—give evidences in favour of this statement.

13. What is the relationship of work with heat ? What is the mechanical equivalent of heat ? What is its value in the C. G. S. system ?

14. A few questions are given below. Put 'Y' against those answers you consider "Yes" and 'N' against those you consider "No".

(a) Is the sun main source of all terrestrial energies ? —

(b) The temperature of a bucketful of boiling water is less than the temperature of a match-stick. But is the heat contained in the water less than that contained in the stick ? —

(c) If you dip your hand in cold water after having dipped in hot water, will you feel the water colder than it actually is ? —

(d) If the temperatures of two bodies are equal, will the heat in them be equal ?

(e) If equal amount of heat be applied to equal masses of different materials, will their temperature change be equal ? —

(f) Does specific heat of a substance depend on (i) the unit of mass and (ii) the unit of temperature ? (i) — (ii) —

(g) Can heat be regarded as a form of energy ? —

CHAPTER FOUR

Light

4·1. Sources of light :

Anything that can give out light may be called a source of light. There are, however, some substances like the sun, the stars, a burning candle, etc., which can themselves emit light. They are known as *luminous* sources. Other sources, which do not themselves emit light but do so on receiving light from luminous sources, are called *non-luminous* sources. The moon is a non-luminous source. It has no light of its own. It reflects sunlight which we call moon-light. Most of the substances that we see around us are non-luminous. Chair, table, trees, buildings and other visible objects receive light from luminous sources and thereby become visible to us.

4·2. Optical medium :

Any medium through which light can travel is an optical medium. If this medium be such that light travels with equal speed in all directions, then the medium is called a *homogeneous* medium. Air, water, glass, etc., are examples of homogeneous optical media.

The homogeneous media through which light can pass very easily, are called *transparent* media. Glass, water, etc., are transparent media.

The media through which light cannot pass, are called *opaque* media. Stone, iron, wood, etc., are opaque media.

Again, the media through which light can pass partly, are called *translucent* media. Ground glass, tissue paper, greased paper, etc., are translucent media.

A ray of light and a beam of light :

In a homogeneous optical medium, light travels in a straight line. So, a straight line may represent the path of a ray of light.

A few such rays together make a beam of light. It is to be borne in mind that it is not possible to produce a single ray of light. However small the source may be, it always produces a beam of light, consisting of several rays.

A beam of light may be of three different types, *viz.* (i) *parallel*, (ii) *divergent* and (iii) *convergent*.

In a parallel beam of light, the rays are parallel to one another [Fig. 1]. Rays of light coming from a very distant source, are regarded as a parallel beam. Rays of sunlight, for example, form a parallel beam. We can, however, form a parallel pencil of rays artificially with the help of a spherical mirror or a lens.

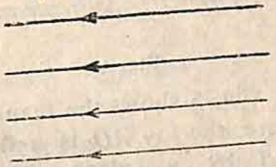


Fig 1.

When rays of light spread out in the form of a cone, so to say, such that the source occupies the vertex of the cone, they form a *diverging* beam of light. (Fig. 2).

On the other hand, if the rays, coming in the form of a cone, converge to a point, they form a *converging* beam (Fig 3).

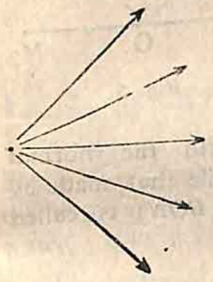


Fig. 2

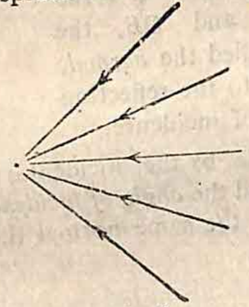


Fig. 8

4·3. Reflection of light at a plane mirror :

We know that in homogeneous media, light travels in straight line. But if light, travelling through one medium, be incident on another, then, a part of the light returns to the first medium from the surface of the second. This phenomenon is known as *reflection of light*. All of you have seen reflection of light by mirrors. In our daily life, we see reflection of light frequently.

If the reflecting surface be very smooth and polished, all the reflected rays will travel in a definite direction, keeping a certain relationship with all the incident rays. In Fig. 4, a beam of parallel rays is incident on a smooth reflecting surface. The reflected rays, also, form a parallel beam. This type of reflection is known as *regular reflection.* As a result of regular reflection, the portion of the reflector on which the rays are incident, appear shining.

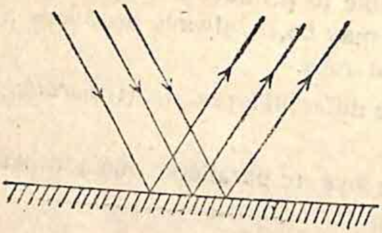


Fig. 4

Fig. 5 shows the regular reflection of a single ray of light. Here, the ray *AO* is reflected along *OB* by a plane mirror M_1M_2. The ray *AO*, called the *incident ray*, is the direction in which light falls on to the reflecting surface. *O* is the *point of incidence* and *OB*, the *reflected ray*. *ON*, called the *normal*, is the perpendicular to the reflecting surface at the point of incidence.

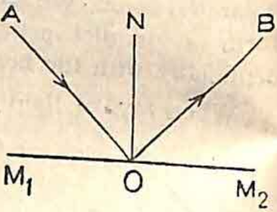


Fig. 5

The angle made by the incident ray with the normal (i.e., $\angle AON$) is called the *angle of incidence* while that made by the reflected ray with the same normal (i.e., $\angle BON$) is called the *angle of reflection.*

Laws of regular reflection :

(1) *The incident ray, the reflected ray and the normal at the point of incidence all lie in the same plane.*

(2) *The angle of incidence is equal to the angle of reflection i.e.,* $\angle AON = \angle BON$.

4·4. Diffuse reflection :

Mirrors or sheets of glass are highly polished and smooth surfaces. The surfaces of most objects, however, are found to be rough and when examined by a microscope, they reveal tiny undulations. When a parallel beam of light falls on such a surface, the individual rays strike it at different angles of incidence. The rays are, therefore, reflected in various directions from the surface (Fig. 6) so that the beam is scattered instead of being reflected as a single beam. We say that the light is diffusely reflected from the surface.

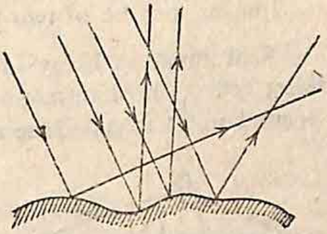


Fig. 6

Ground glass, white paper, wall, cinema screens, etc., being rough, produce diffuse reflection of light. One advantage of this is that these substances look equally bright from whatever direction they are viewed. But, plane mirror and other smooth surfaces produce regular reflection ; hence the portion of those surfaces on which light falls, looks very dazzling.

4·5. Image and its definition :

All of you have seen images. When we stand before a looking glass, we see images of our appearance. The image of a tree, standing by the side of a pond, is seen reflected on the sheet of water. How is this image produced ?

When rays of light from an object reach our eyes directly, we see the object proper. But when we look into a plane mirror we seem to see objects which are really in front of the mirror, in positions behind the mirror. When we look through a telescope, stars seem to be in positions nearer to us than they really are, and when we use a magnifying glass, the object we are examining seems to be further away than its true position. Whenever an object seems to be in some position differing from its true position, it is said that an image of it exists at its apparent position.

The formation of an image is always due to the fact that, on their journey from the object to the eye, the rays of light are changed in direction, due to reflection or refraction, so that they appear to come from points other than those from which they started.

Definition : When rays of light coming from a point source, after reflection or refraction, actually meet at another point or appear to diverge from another point, the second point is called the image of the first point.

Images may be of two types, *viz.* (i) *real* and (ii) *virtual.*

Real image : If rays of light coming from a point source, after reflection or refraction, actually meet at another point, the second point is called the real image of the first point. In Fig. 7, rays of light from the point P after being refracted by a convex lens L, actually meet at the point P'. If the object P is viewed through the lens, the eye will see the image P' instead. This image is a real image. If a screen be placed at the position P', the image of the point will be cast on the screen.

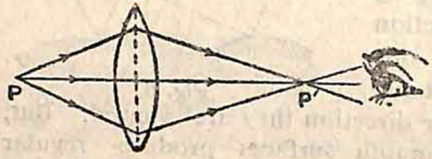


Fig.7

Virtual image : If rays of light coming from a point source, after reflection or refraction, appear to diverge from another point, the second point is called the virtual image of the first point. Virtual image cannot be received on a screen. The image we see in a looking glass is a virtual image. If a screen be placed at the back of the looking glass where the image seems to be seen, we will find no image on the screen.

Plane mirror produces virtual image :

Suppox a candle stands in front of a plane mirror In Fig. 8, P is a point object on the candle. Rays of light are coming out of the point P and after reflection at the plane mirror M_1M_2, reach the eye of an observer. The eye cannot follow the bent direction of the rays. Looking through the mirror, the eye will see, as if, the rays are coming from the point P' at the back of the mirror. In other words, P' is the virtual image of the point P. In this way a complete image of the

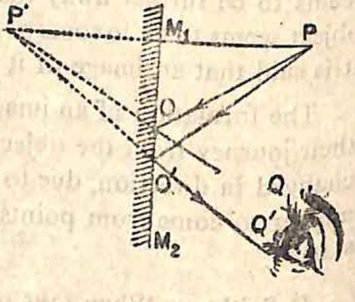


Fig. 8

whole of the candle will be seen through the mirror. If a screen be placed at the position P', no image will be seen on the screen.

It is to be noticed that *the image is formed at the same distance behind the mirror as the object is in front.*

Many such examples of virtual image may be cited. The image of a tree cast on a sheet of water is a virtual image. The sheet of water, here, serves the purpose of a plane mirror.

Differences between a real and a virtual image :

(1) If the rays, coming from a point, after reflection or refraction, actually meet at another point, a real image is formed. On the other hand, if the rays, instead of actually meeting, appear to diverge from another point, a virtual image is formed.

(2) A real image can be seen and can be cast on a screen as well. But a virtual image can be seen only and cannot be cast on a screen.

Lateral inversion of an image in a plane mirror :

When we stand before a looking glass our right hand seems to be the left hand of the image and vice versa. It will be obvious to you if you raise your right hand to your right ear, the image will be seen to have raised its left hand to the left ear. Write a letter, say "R" on a piece of paper and hold it before a mirror. You will see that the image is inverted laterally (Fig. 9). This change in the image is known as *lateral inversion*. No lateral inversion is, however, seen in the images of objects which are symmetrical.

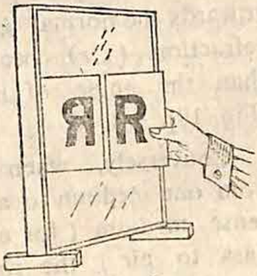


Fig. 9

4·6. Refraction of light at a plane surface :

Look at the bottom of a vessel full of water. It will appear much shallower than it actually is. A straight stick appears bent when partly immersed in water. These and many similar effects show that light travels in a different direction in water than it does in air. In other words, there is a change in direction of light when it passes from one medium to another. This change in direction of light is known as *refraction*.

Suppose, a ray of light, travelling along the straight line AB in air, is incident obliquely on a block of glass [Fig. 10]. The ray will now enter into glass block. But the direction the ray takes up in the glass block is different from AB, because refraction of light will take place at B. Suppose, the ray travels in the glass along the straight line BC.

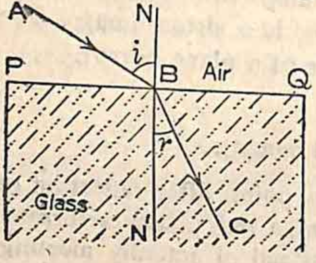


Fig. 10

Here, AB is called the *incident ray*, BC, the *refracted ray*, B, the *point of incidence* and NN', *the normal* at the point of incidence to the surface of separation PQ between the two media, *viz*., air and glass. The angle between the incident ray AB and the normal BN (i.e., $\angle ABN$) is called the *angle of incidence* and that between the refracted ray BC and the normal BN' (i.e., $\angle CBN'$), the *angle of refraction.*

It is important to remember that, when a ray passes from one medium to a more optically dense medium (for example, from air to glass) the ray bends towards the normal *i.e.*, the angle of refraction ($\angle r$) becomes smaller than the angle of incidence ($\angle i$) [Fig. 10].

Conversely, when a ray passes from one medium to a less optically dense medium (for example, from glass to air) the ray bends away from the normal *i.e.*, the angle of incidence ($\angle i$) becomes smaller than the angle of refraction ($\angle r$) [Fig. 11].

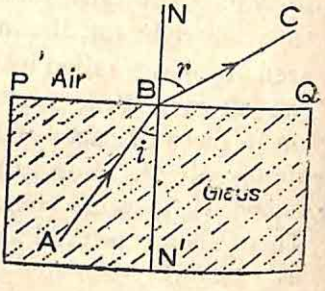


Fig. 11

4·7. Experiments on refraction of light through water and glass :

(A) *Refraction through water* : Pour water in a jar of average size and place a thick, black cardboard by the side of the jar. Make a small hole on the cardboard so that a fine beam of sun-light may pass through the hole and may fall on the water surface. It will be seen clearly that light travels along a different direction in the water and has produced a circular patch on the bottom of the jar [Fig. 12]. If, now, water be withdrawn from the jar, it will be seen that the patch has been formed on the wall of the jar and it is in the same straight line with the hole on the cardboard.

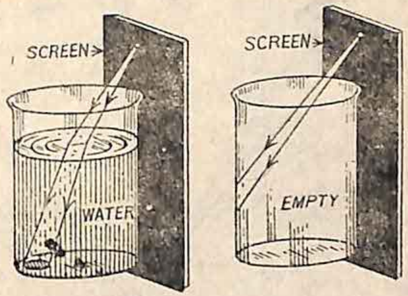


Fig. 12

This experiment, therefore, shows that light, while passing from air to water, has changed its direction.

(B) *Refraction through glass* : Put a dot mark on a piece of paper and place a slab of glass over it. Now, try to look at the dot from above. It will appear to be nearer to the eye than it actually is. This is due to refraction of light.

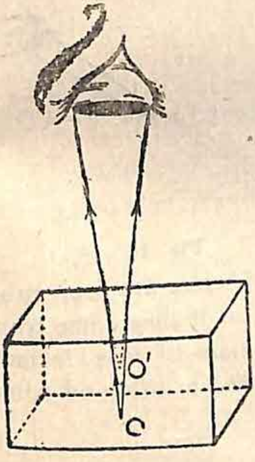


Fig. 13

Beams of light travelling upwards from the point O [Fig. 13] at the bottom of the glass block bend away from the normals at the upper surface and so enter the eye, as if, they come from O' so that an image of O is formed at O'.

For similar reasons, water in a tank viewed from above appears less deep than it actually is, as you may have observed at the swimming baths.

4·8. Some effects of refraction :

(A) *Bending of a stick immersed in water* : The apparent upward bending of stick when immersed obliquely in water is shown in Fig. 14. Rays of light from the end *B* of the stick pass from water to air and are bent away from the normal because they are passing to a less optically dense medium. Reaching the eye, the rays appear to come from a point *A* above *B*. *A* is thus the image of *B* as a result of refraction. The same reasoning applies to any point on the immersed portion of the stick so that the observer sees an image apparently in the bent position.

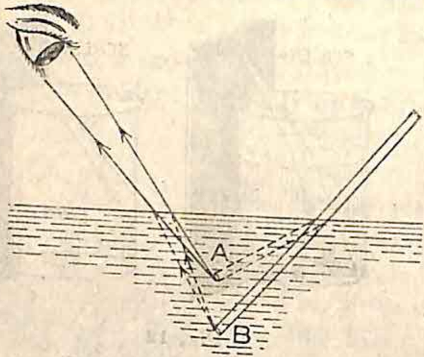


Fig. 14

(B) *The image of a coin dipped in water* : Place a coin in a porcelain basin and place the eye in such a manner that the coin is just out of view. In this case, the rays from the coin will be obstructed by the edge of the basin and will not be able to reach the eye. Keeping the eye in that position, pour water in the basin. The coin will, at once, come into view. This is due to refraction of light. Beams of light from the coin incident on the surface of water are bent away from the normals and so enter the eye, as if they come from a point *P'* above the point *P*, so that an image of *P* is formed at *P'*. The coin, thus, will appear to be raised a little and come into view again.

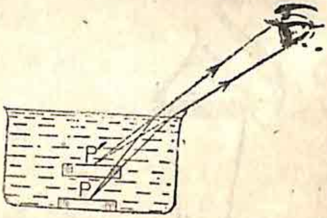


Fig. 15

(C) *Multiple images in a thick plane mirror* : Ordinary mirrors made by silvering the back face of a sheet of glass, suffer from the disadvantage that extra images are formed by reflection from the front surface of the glass and also by multiple internal reflections inside the glass. Fig. 16 shows how multiple images of a candle are formed by a thick mirror.

From the point P of a candle, a ray of light PA is incident at the point A on the mirror. The dark line ACD shows the path followed by most of the light in the beam PA, giving rise to the main image Q_2. A little of the beam is reflected at A instead of entering the glass, and gives rise to the image Q_1 in front of Q_2. Some of the light is reflected at D instead of leaving the glass, and after reflection at the back surface once more, reaches G. There, part of it leaves the glass, giving rise to a faint image at Q_3 but part is reflected, and so on. In this way, multiple images are formed in a thick plane mirror.

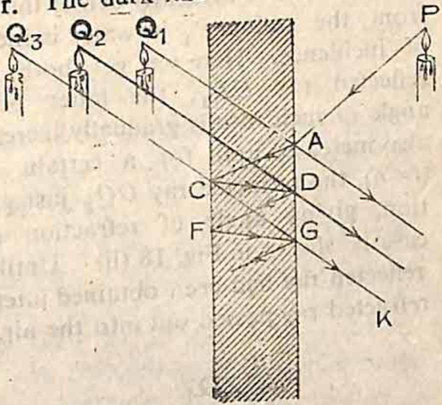


Fig. 16

4·9. Total internal reflection :

When light passes from one medium to a more optically dense medium, there will always be both reflection and refraction for all angles of incidence at the surface of separation between the media. In Fig. 17, a ray of light, coming from air enters a glass block and finally emerges into air again. In both the cases, reflection and refraction have taken place. But this is not always the case if light passes from one medium to a less optically dense medium.

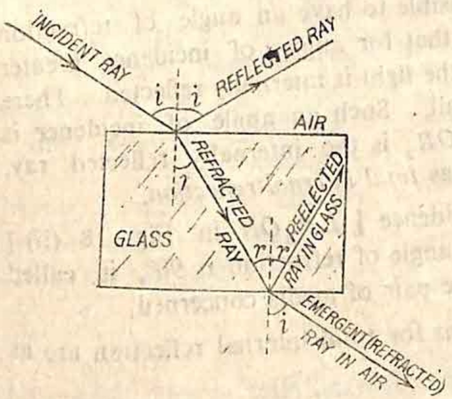
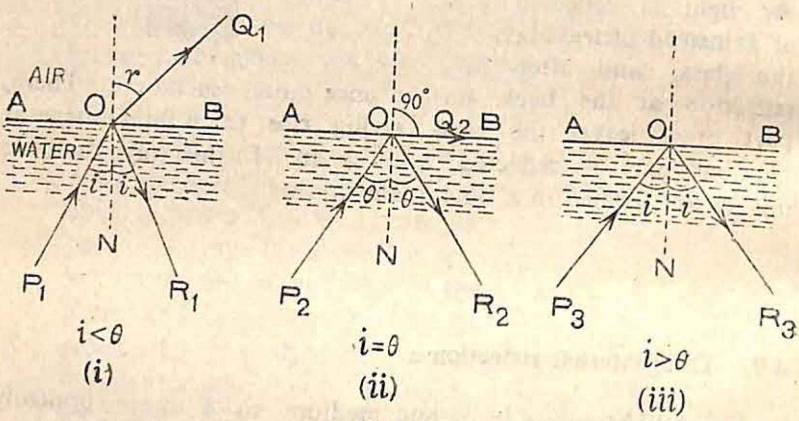


Fig. 17

Suppose AB is the surface of separation between air and water. Water is evidently denser than air. A ray P_1O [Fig. 18(i)] from the point P_1 in water is incident at O at a small angle of incidence. Here we get both a refracted ray OQ_1 and a reflected ray OR_1, the latter being relatively weak. If the angle of incidence is gradually increased, the angle of refraction also increases until for a certain critical angle of incidence ($i=\theta$), the refracted ray OQ_2 just grazes the surface of separation, giving angle of refraction equal to 90°. This special case is shown in Fig. 18 (ii). Until now, a comparatively weak reflected ray had been obtained internally in water, and a strong refracted ray passed out into the air.



$i<\theta$
(i)

$i=\theta$
(ii)

$i>\theta$
(iii)

Fig. 18

Since it is clearly impossible to have an angle of refraction greater than 90°, it follows that for angles of incidence greater than the critical angle, all the light is internally reflected. There is now no refracted ray at all. Such an angle of incidence is shown in Fig. 18(iii) and OR_3 is the internally reflected ray. This condition is described as *total internal reflection*.

Besides, the angle of incidence [$\angle P_2ON$ in Fig. 18 (ii)] corresponding to which the angle of refraction is 90°, is called the *critical angle* between the pair of media concerned.

Thus, *necessary conditions* for total internal reflection are as follows :—

(1) The ray of light must travel from one medium to a less optically dense medium.

(2) The angle of incidence must be greater than the critical angle between the pair of media concerned.

4·10. Some effects of total internal reflection :

(1) Put an iron ball painted black with lampblack in water. The ball will appear silvery although painted black. This is due to total internal reflection of light.

Because of the coating of lampblack a layer of air will stick to the ball even when it is immersed in water. Rays of light, in going through water, are incident on that air-layer and if the angle of incidence is greater than the critical angle for water and air, the rays will be totally reflected upwards. Reaching eye, the rays will give a silvery appearance to the ball.

For the same reason, when bubbles come up from a certain depth of water, they appear shining. Diamonds and other valuable stones owe their brilliance to the total reflection of light.

(2) Place a test-tube partly filled with water, slantingly in a beaker containing water as shown in Fig. 19. If the empty portion of the test-tube is viewed from above, it will have a silvery and mirror-like appearance.

The rays of light, in going from water to air in the test-tube, get totally reflected and reach the eye. If the portion of the test-tube filled with water be similarly viewed, it will not appear silvery because rays of light will pass on straight through the water.

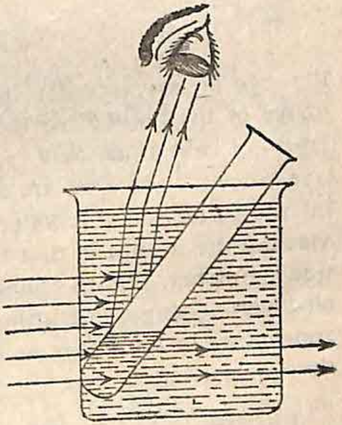


Fig. 19

4·11. Application of total internal reflection in nature :

In deserts and in cold countries people may suffer from an optical illusion with respect to distant objects. The traveller in a desert often sees what appears to be a sheet of water a short distance ahead of him. This he is never able to reach, since it is an optical illusion. In cold countries, it is sometimes seen that an inverted image of a distant object is hanging in the sky. This optical illusion is known as **mirage**. It is brought about by total reflection of light.

Mirage in a desert : The layer of air in contact with the hot sand-bed of the desert becomes very hot and expands, thus becoming optically less dense than the cooler layer above. A ray of light from the point P, say, of a distant tree, coming through the cool air-layers bends away from the normal slightly as it enters the layer of warm air of decreasing density nearer the sand-bed. The light thus bends away from the normal progressively as it traverses the lower layers until it meets a hot layer, at Q, say, near the bed, at an angle of incidence greater than the critical angle. Total reflection then occurs and the light enters the eye of the traveller as shown in the Fig. 20. A virtual image of the point P is, therefore, seen at P'. The hot surface layer of air thus acts as a mirror in which an image of the tree is seen. Since we are accustomed to seeing a tree mirrored in the surface of still water, the natural assumption made when viewing the mirage is that there is a pool of water near the tree. Further, due to change of temperature, the density of the air-layer changes continuously. This gives a quivering appearance to the image as if the water-surface is disturbed by slow air-currents.

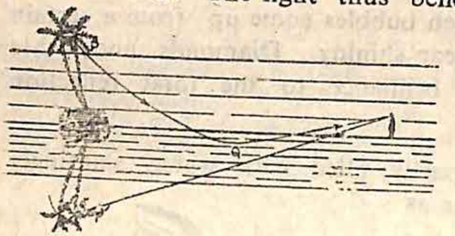


Fig. 20

For the same reason, on a hot summer day, the surface of a hot asphalted roadway in the near distance may appear wet and shiny as though it had been raining.

Mirage in cold countries : In cold countries, the upward layers of air are gradually less dense while the layer near the ground denser. Beams of light travelling upwards pass from one layer to a less optically dense layer and are, therefore, bent away from the normal. Finally the ray is totally reflected downwards at some particular layer where the angle of incidence of the ray exceeds the critical angle. When the beams of light reach the eye of an observer, they produce an inverted image of the object in the air (Fig. 21).

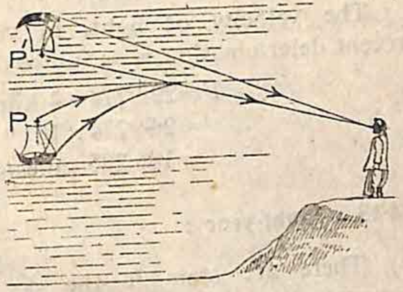


Fig. 21

4·12. Velocity of light :

It has been found that light travels, in air, with a speed of about 186,000 miles per second. So, the velocity of light is enormous ! As a matter of fact, ultimate speed-limit of material bodies is the velocity of light.

The sun is at a distance of about 93,000,000 miles from the earth. Light, with the above tremendous speed, takes about 8·3 minutes to reach the earth from the sun. But there are stars and planets in the sky which are further and further away from the earth than the sun. Consequently light from those heavenly objects will take greater time to reach the earth.

If any incident is noticed on those stars from the earth at any instant, it is to be remembered that the incident happened long before that instant. For example, the light that reaches the earth to-day from the nearest fixed star 'Alpha centuri', had started from that star 4·4 years ago ! If the brightest star 'Sirius' in the space suddenly dies to-day, it will continue to send light to the earth for about 8·8 years more. In other words, the star will be visible from the earth for 8·8 years more. It may, therefore, be said that many of the stars that we see to-day in the sky, may have disappeared long ago !

The velocity of light was first experimentally determined by a Danish astronomer, Romer. Afterwards, Fizeau, Michaelson, Anderson and others carried out experiments for the determination of velocity of light.

The velocity of light, in vacuum, according to the most recent determination is as follows :

$$\begin{aligned} V &= 299{,}774 \pm 4 \text{ kilometres/sec.} \\ &= 2{\cdot}99774 \times 10^{10} \text{ cm./sec.} \\ &= 186{,}285 \text{ miles/sec.} \end{aligned}$$

4·13. Light-year :

There are stars in the sky which are so far apart, that gigantic figures will be required to express their distance, if mile is taken as unit of distance. To express these astronomical distances, light-year is taken at the unit.

Definition : The distance travelled by light, having a velocity of 186,000 miles per sec., in one year is called a light-year. So,

$$\begin{aligned} 1 \text{ light-year} &= 186{,}000 \times 365 \times 24 \times 60 \times 60 \text{ miles.} \\ &= 5{\cdot}86 \times 10^{12} \text{ miles (nearly)} \end{aligned}$$

Alternatively : $1 \text{ light-year} = 300{,}000 \times 365 \times 24 \times 60 \times 60$ kilometres.

$$= 9{\cdot}45 \times 10^{12} \text{ kilometres (nearly)}$$

4·14. Lenses and refraction of light :

Lenses have been in use for centuries and were well known to the Greeks and medieval Arabs. Utilising the converging action of a lens, *burning glass* was devised many centuries ago. Lenses play an important part in our own daily life. Apart from the benefit of spectacles which enable millions of people to see things with comfort, our lives would be vastly changed if we had no camera, microscopes, telescopes and other optical instruments, all of which employ lenses.

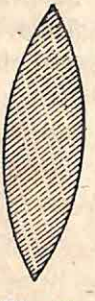


Fig. 22

Any transparent refracting medium bounded by two spherical surfaces or one spherical and other plane surface, may be called a lens.

If the lens is thicker at its centre than at its edges, it is a convex or converging lens. (Fig. 22)

Focussing action of a convex lens :

Let us take a convex lens and suppose a ray *PB* from the point *P* of an object-point is incident on the lens at the point *B* [Fig. 23]. Since the ray is passing from air into glass, the ray will bend towards the normal at *B* and will travel along *BC*, because glass is optically denser than air. When the ray will emerge into air again at *C*, it will move away from the normal at *C* and will travel along *CQ*. Another ray *PF* from the same point *P*, being incident at a smaller angle of incidence at *F* than before will bend less and emerge into air again at the point *G* after travelling along the path *FG* inside the lens. It will be seen that this ray also passes through the point *Q*. Similarly another ray *PM*, incident at the centre of the lens, will pass on without deviation, because it is incident normally on the lens. It will, also pass through the point *Q*. Since the lens is bounded by two spherical surfaces all the rays coming from the point *P* will be likewise, found to pass through the same point *Q* after refraction through the lens. This action of a convex lens is called its *focussing action*.

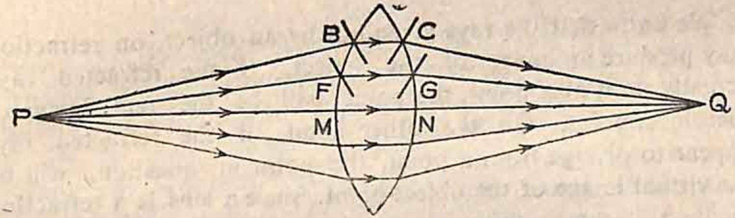


Fig. 23

Focus and Focal length :

If a beam of parallel rays, travelling parallel to the principal axis of a convex lens, are refracted by the lens, the rays become converging and intersect each other at a particular point on the axis [Fig. 23 (*a*)]. The point is called the *focus* of the convex lens. In the figure, *F* is the focus of the lens.

Generally, we use lenses where surfaces have equal curvature. In such lenses, if we take a point on the principal axis inside the lens equidistant from the two surfaces, the point is called the *optical centre* (O) of the lens. If the lens be thin, then a ray directed towards the optical centre, suffers no refraction and passes out straight through the lens.

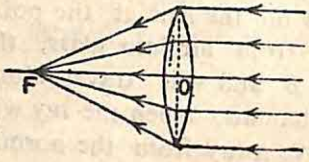


Fig. 23 (a)

The *focal length* of a lens is the distance between the optical centre (O) and the focus (F) of the lens.

Formation of images by lenses :

We know that the rays given out by an object, on refraction may produce an image of the object. If the refracted rays actually meet at a point, the point will be the real image of the object-point. On the other hand, if the refracted rays appear to diverge from a point, the point in question, will be the virtual image of the object-point. Since a lens is a refracting medium, it can produce, in the above process, both real and virtual images of an object.

Experiments : Place a burning candle and a white paper-screen fixed in a suitable stand some distance apart. Now fix a convex lens in a lens-holder and place it in between the candle and the screen. See that the tip of the candle flame, the centre of the lens and the centre of the screen are all in one straight line. Now move the screen until a clear but inverted image of the candle flame is formed on the screen [Fig. 24]. In this case, the image is real. It is to be remembered that real images are always inverted.

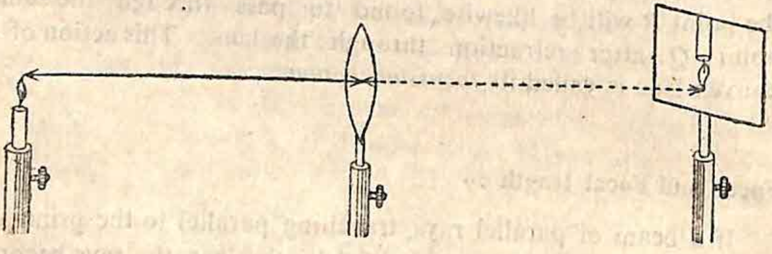
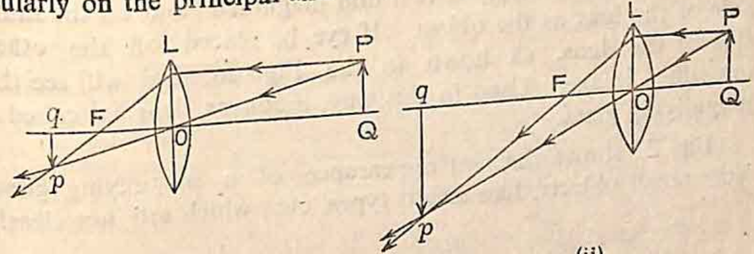


Fig. 24

4-15. Construction of ray diagrams of images :

Geometrical construction of images formed by a convex lens may be done in the following way :

LO is a convex lens and *PQ* is an object standing perpendicularly on the principal axis of the lens at a sufficient distance away from the lens. A ray *PL* [Fig. 25(i)] parallel to the principal axis passes through the focus *F* after refraction through the lens and takes up the direction *Fp*. Another ray *PO*, passing through the optical centre of the lens, will come out straight without any deviation. Since the two refracted rays actually meet at a point *p*, it will be the real image of *P*. A perpendicular *pq* drawn on the axis, will be the real image of *PQ*. If the object *PQ* is far away, the image *pq* will be very small in size and will be inverted.



(i) (ii)

Fig. 25

On the other hand, if the object *PQ* be situated just on the right of the focus, its image will be greater than the size of the object. The geometrical construction of the image is illustrated in fig. 25 (ii).

4-16. Convex lens as a magnifying glass :

Suppose *PQ* is a very small object to be magnified by a magnifying glass [Fig. 26]. A convex lens is so placed that the object *PQ* falls *within its focal length*. A ray *PL* from *P*, parallel

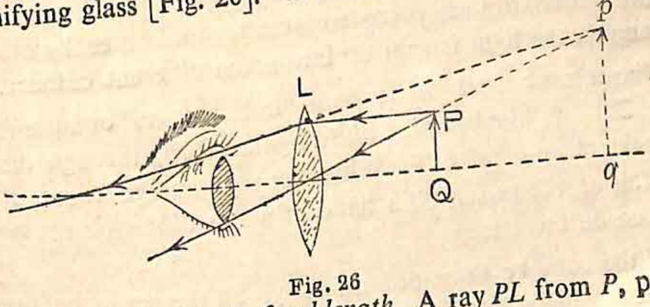


Fig. 26

to the principal axis passes through the focus after refraction through the lens. Another ray from *P*, passing through the optical centre of the lens, will come out straight without any deviation. The two refracted rays appear to diverge from a point *p* and *pq* is the image of the object. It will be noticed that the image is erect, virtual and magnified and on the same side of the lens as the object. If eye be placed on the other side of the lens, as shown in the Fig. 26, eye will see the magnified image. Used in this way, a convex lens is called a magnifying glass.

Fig. 27 shows the real appearance of a magnifying glass. Very small objects, like small types, etc., which are not clearly visible with naked eye, may be seen distinctly by a magnifying glass. A round metallic frame holds the lens and a handle is provided to use it properly.

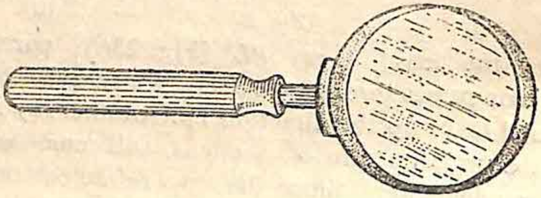


Fig. 27

4·17. Dispersion light :

Dispersion of light was first discovered by Sir Isaac Newton in 1666. He observed that if white light coming from the sun be allowed to pass through a glass prism (a glass block, two of whose boundaries are plane surfaces inclined to each other at an angle), the light is split up into seven different colours.

Experiment : *H* is a small hole in an opaque screen [Fig. 28]. A fine beam of white light, coming through the hole, in incident on a prism *P*. When the rays emerge from the prism and fall on the screen *S*, a band of seven different colours will be seen on the screen.

If the band be examined carefully, all the seven colours of a rainbow will be seen in it. At one edge of the band, there is *red* and at the other *violet*. The other colours from red end are *orange, yellow, green, blue* and *indigo*. The sequence of colours

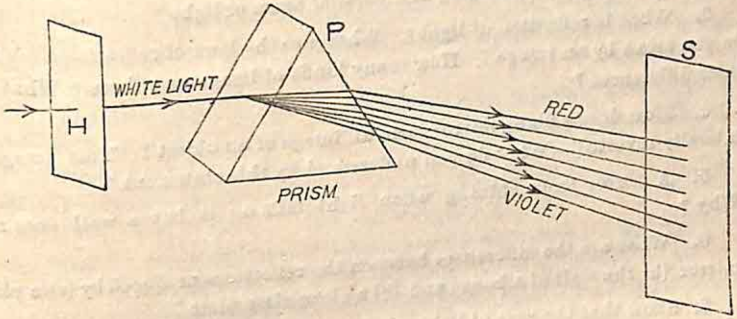


Fig. 28

can be best remembered by the word 'VIBGYOR' which is formed by taking the initial letters of each colour.

The band is known as *spectrum*. The phenomenon of splitting up of white light into the constituent colours due to the passage through a prism is called *dispersion of light*.

When we say that the spectrum is made up of the colours red, orange, yellow, green, blue and violet, we must remember that these divisions are made merely for convenience. The whole spectrum really consists of an endless number of colours. Experts can distinguish thousands of shades. The eye also recognises colours that are not found in the spectrum, but which are effects produced by mixing various spectrum colours. *Purple*, for instance, is a mixture of blue and red ; *pink* is a mixture of red and white ; *browns* are produced by mixing red and yellow. It has been found by experience that virtually any recognisable colour can be produced by mixing different proportions of the three colours red, green and blue-violet. These may be called the *primary* colours. If three such spots of light are thrown on a screen, the place where all three overlap will appear white.

Exercise

1. What is an optical medium? What are transparent, translucent and opaque media?

2. What is the difference between a ray of light and a beam of light? What are converging, diverging and parallel beam of light?

3. What is reflection of light? What are the laws of reflection? What do you mean by an image? How many kinds of images are there? What are their differences?

4. How does a plane mirror form an image of an object? 'This image is laterally inverted'—What do you understand by this statement?

5. A mirror looks shining when light falls on it but a wall does not. Why?

6. What are the differences between the reflections produced by (a) a plane mirror (b) the wall of a house and (c) a clean glass plate?

7. Show that the rays of light coming out of a point source, after reflection at a plane mirror, appear to diverge from a second point. What is that point called? Where is it situated? What is its nature?

8. What is refraction of light? Explain, with neat diagrams, how refraction takes place in the following cases:—(a) From air to glass (b) From water to air.

9. Answer the following questions:—

(i) A stick immersed slantingly in water looks broken. Why?
(ii) A vessel full of water looks less deep when looked from above. Why?
(iii) A thick mirror produces multiple images of an object. Why?

10. Explain fully what you understand by total internal reflection and critical angle. State whether critical angle will be available in the following cases:—(i) Light is passing from air into glass, (ii) Light is passing from glass into air.

11. Answer the following questions:—

(a) A metallic ball coated with lampblack looks silvery when immersed in water. Why?

(b) An empty test-tube immersed obliquely in a vessel containing water, looks silvery. Why?

(c) In hot summer days, an asphalted road looks wet at certain distance away although there is no water there actually. Why?

12. What is a mirage? Explain with a neat diagram how it is produced?

13. What is the velocity of light? 'Some of the stars that we see in the night sky, may have disappeared long ago.'—Explain the statement.

14. What is a light-year? How many miles make a light-year?

15. What is a convex lens? Define its focus and focal length. Is light reflected or refracted through a lens?

16. Explain, with a diagram, the magnifying action of a convex lens.

17. What is dispersion of light? What is a spectrum? What are the constituent colours of white light?

18. A few questions, along with their probable answers, are given below. Put a √ sign against those answers which you consider most satisfactory :—

(i) What is that medium called, through which light can partly pass? *Ans.* transparent, translucent, opaque.

(ii) Why is the cinema screen made white? *Ans.* for regular reflection; black pictures appear distinct against white background; for diffuse reflection.

(iii) Why do multiple images appear in a thick mirror? *Ans.* due to multiple reflection; due to multiple refractions; due to multiple and simultaneous reflections and refractions.

(iv) What are the characteristics of an image produced by a plane mirror? *Ans.* equal size; real; inverted; erect and of double size.

(v) What are the conditions of total internal reflection? *Ans.* Light should pass from one medium to more optically dense medium; angle of incidence should be greater than the critical angle; light should pass from one medium to less optically dense medium.

(vi) What type of image a convex lens can produce? *Ans.* real; magnified; virtual; reduced; of equal size.

CHEMISTRY

CHAPTER ONE

Matter and Materials

Nature has laws of its own, deviation from it is yet to be found. Hydrogen and oxygen combine together to form water. This is always true ; if appropriate conditions could be produced, water is formed and any person willing to seek the truth of it can do the same with the help of suitable apparatus. These facts again are of various types. A hormone, for example, can turn a normal man into a lunatic one. Chloromycetin kills typhoid bacteria. An animal, as distinct from plant, cannot prepare food in its own body directly from nature. If a lodestone is dipped into iron filings tufts of them will cling to its poles etc. etc.. With the accumulation of facts of such diverse nature, connected theories, various laws, etc, science has been divided into different branches more or less for the sake of convenience. One such branch of science is called Chemistry and it is, therefore, a small fraction of the sum-total of human knowledge. But science, a coherent whole, that extends over the entire realm of nature cannot be marked off into sharply defined regions. With the progress of time the boundaries between such regions have become blurred.

The word 'chemeia' has its origin in 'qemi', the ancient name of Egypt meaning the Black land. The term 'chemistry' is probably derived from the word 'chemeia'. In material world we come across a wide variety of substances ; water, glass, air, sand, chalk, iron and sulphur are some of the examples. It is evident that all these are distinctly different bearing even no external resemblance to one another. Each substance is distinguished by its *properties* ; 'properties' applied to substances refer to their characteristic qualities. Water, for example, is a colourless, odourless liquid, freezes at 0°C and boils at 100°C under normal atmospheric pressure. Properties are of two types —physical and chemical. Properties that can be tested without changing the nature or composition of the substance under investigation are termed physical properties. The properties of water as stated above are physical properties. But the properties which can be tested only by changing the nature or composition of the substance are termed chemical properties. Thus magnesium ribbon when burns in air, white fumes evolve. These fumes get deposited on a cold surface as a white solid. This white solid is what is known as magnesium oxide, differing qualitatively from the original substance magnesium. From this it can be said that magnesium is combustible—the ability to burn is a chemical property, because it can be tested only by changing the chemical nature of the substance concerned. Again when coal or wood burns, iron rusts, new substances are formed. In each case a new substance or substances with new properties are formed. Such a phenomenon accompanied by a radical change of the substance or substances is called a chemical phenomenon and the science that primarily deals with these phenomena, the laws that govern them and the effects of heat, pressure, etc. on these changes is called *chemistry*.

Aims of chemistry are : (i) To get acquainted with the properties of matter through experiments and observations, (ii) to ascertain the composition of matter or bodies (to discover the relations existing among the components), (iii) to know the changes (if any) in composition under the changing external conditions such as pressure or temperature, or when the substances mutually react upon one another, (iv) to produce (if possible) the original substance artificially starting from its very constituents and even, (v) to produce compounds non-existent in the works of nature.

1-1. Matter. Bodies or objects that we come across in our everyday life are composed of matter. Anything that possesses mass, occupies space, offers resistance and can be perceived through one or more of our senses is called matter.

Matter can exist in three states : solid, liquid and gas.

A *solid* has a definite shape and volume of its own. A piece of chalk wlll always have the same shape and will always occupy the same space.

A liquid possesses a definite volume but has no definite shape of its own. It assumes the shape of the vessel in which it is put. If a liquid from one vessel be poured into another, it takes the shape of the latter container but no change in volume occurs *i.e.*, the original volume is retained. Oil, water, mercury, etc. are some of the well known liquids.

A gas has no definite shape or volume. Thus a small quantity of a gas can fill both a big and a small vessel *i.e.*, it may fill any container completely and to a uniform density irrespective of the capacity of the latter. Hydrogen, oxygen, nitrogen, etc. are gaseous substances.

In spite of the above differences all kinds of matter have a common property, the proporty of having a mass. But weight, it is to be noted, though a general property, is not an essential property of matter. When the position of a body with respect to the earth changes, the weight changes.

The smallest particle of any form of matter which can exist in free state possessing all the properties of the substance is called a molecule, or in other words all matters are composed of molecules. These molecules are separated from one another by intermolecular space. When a body is subjected to pressure or its temperature is lowered, contraction in volume occurs. Conversely if the pressure is reduced or the temperature is raised, the volume increases. Again, the formation of solution in some cases results in contraction of volume. When some sugar is mixed with water, the former disappears ; the sweet taste of the mixture indicates the presence of sugar in it. Though the sugar has thus undergone solution corresponding increase in the volume of water is not observed. All these facts can be explained by assuming the existence of spaces in between molecules.

The difference in three states of matter lies in the closeness with which the constituent molecules are packed.

In a substance which is nothing but an assembly of molecules, the molecules are in a state of motion and are subjected to an intermolecular force of attraction. In a gas the intermolecular force of attraction is almost negligible, the molecules are widely separated from each other and are capable of moving in any direction independent of each other. Hence a gas cannot preserve any definite shape or volume. In a liquid the intermolecular force of attraction is comparatively great, the intermolecular space is smaller. The molecules are packed more closely than in a gas. Hence its volume is limited by definite bounding surface. A liquid thus preserves a definite volume. But nevertheless some freedom of movement being there a liquid, like a gas, has no definite shape, and so takes the shape of the vessel in which it is put. In a solid the force of attraction is still increasingly manifested, the intermolecular space is minimum and the result is that a solid acquires the ability to preserve a definite shape and volume of its own. Hence a solid has a fixed pattern.

Application of high pressure and low temperature results in a decrease of intermolecular space (and velocities). Conversely if the pressure is decreased and the temperature is raised, intermolecular space increases. No physical state in general is, therefore, permanent for a substance. Under suitable conditions created by a careful control of pressure and temperature a substance may be made to pass from one state to another. With ordinary conditions of pressure and temperature it exists in one state only. This explains why a solid generally passes on to a liquid and the latter into a gas on heating ; and again why the same changes in the reverse order takes place on cooling.

A pure solid is characterised by its melting point, while a pure liquid has its characteristic boiling point. This has already been discussed in the chapter on 'change of state'.

1-2. Physical and chemical changes :

What happens, if some water is kept in an open shallow vessel or if a piece of iron is kept exposed in moist air ?

From our everyday experience we can answer the questions easily. Water in the open vessel slowly evaporates ; the piece of iron rusts. In each case the material undergoes a change. In fact if we look around us, we see that numerous changes are going on constantly. Some changes are spontaneous and some again are brought about by creating necessary conditions artificially. The changes that matter undergoes are of two types : physical and chemical.

A change, which does not involve any change in composition of the substance is called a *physical change.*

Water on heating changes to steam, which on cooling condenses and the same water is obtained. Cooling freezes water into ice and if ice is allowed to melt water is obtained again. This involves no change in the composition of the material water ; steam, water and ice are the different states of the same substance. Hence, the changes are physical changes.

A platinum wire on heating first becomes red hot and then white hot. In this condition it emits light. On cooling the original colour of the wire is restored and it loses the property of emitting light. This involves no change in the composition of the material platinum. The change is thus a physical one.

If a little common salt is shaken with water, the former disappears. On evaporating away water, the common salt is obtained unchanged in mass and composition. The change is undoubtedly a physical one.

We need not multiply the examples. Let us now try to define clearly the term chemical change in which a chemist is primarily interested. A change that alters the composition of a substance or substances taking part in the change is termed a *chemical change.*

When a piece of iron is left to itself in moist air, a brown crust is formed on its surface. This phenomenon as has already been stated is known as rusting of iron. The brown crust is called rust. The composition and properties of rust are different from those of iron itself. There has been a radical change, a new substance with entirely new properties is produced. The rusted iron weighs more than the original iron. Roughly speaking the iron here is converted into iron oxide, compound of iron. Thus, it is a chemical change.

When an electric current is passed through fused common salt, a yellowish green gas known as chlorine and a metal called sodium are obtained. The properties of common salt are entirely different from those of chlorine and sodium. Thus the change is decidedly a chemical one.

A plant, in presence of sunlight and with the help of chlorophyll, the green colouring matter in its leaf synthesises its carbohydrate food sugar from carbon dioxide and water in the laboratory of its own body. Both carbon dioxide and water vapour required in the process are quite different from the resultant sweet-tasting sugar. The change is undoubtedly a chemical one. Energy required in the process is obtained from sunlight.

When water is sprayed on a lump of quick lime, much heat is evolved and the lump crumbles to powder with a hissing sound. Water here combines with quick lime and changes it to slaked lime. Quick lime and slaked lime are quite different substances. The change is permanent and it is a chemical one.

Characteristics of physical and chemical changes

Physical change	*Chemical change*
(i) No new substance is formed, composition remains unchanged.	(i) Involves a change in composition, new substances with new properties are formed.
(ii) The change is temporary ; if the cause of the change is removed, the substance usually returns to its original form.	(ii) The change is permanent ; new substances cannot be easily reverted to the original conditions.
(iii) Thermal change may or may not take place.	(iii) Chemical change is always attended either with evolution or absorption of heat.
(iv) No change *i.e.*, loss or gain in weight occurs.	(iv) The material undergoing the change usually suffers a loss or gain in weight.

1-3. Endothermic and exothermic reactions

You know that in a chemical change heat is either absorbed or evolved. The reactions in which heat is absorbed are known as *endothermic* reactions, while chemical reaction which evolve heat are called *exothermic.* The compounds formed from their elements with absorption of heat are called *endothermic* compounds, whilst those formed from their elements with evolution of heat are called *exothermic* compounds.

In the combination of nitrogen and oxygen to form nitric oxide, heat is absorbed. Hence the reaction is endothermic and nitric oxide is an endothermic compound. When coal burns in air, heat evolves. Carbon of coal combines with oxygen of the air to form carbon dioxide. The reaction is, therefore, an exothermic one and the product carbon dioxide is an exothermic compound. Many other examples of these types may be cited.

1-4. To initiate and to influence a reaction :

Another important aspect of chemical changes that deserves consideration here is the various ways of initiating reactions and the creation of appropriate conditions for their facile passage.

No reaction between white phosphorus and iodine takes place, if they are carefully kept apart in a basin. But phosphorus bursts into flame immediately on coming in contact with iodine and phosphorus tri-iodide is formed. Chemical change here has been brought about by actual *contact* of the reactants.

When dry, no reaction occurs between sodium bicarbonate and tartaric acid even on thorough mixing. But if a little water is added to the mixture, reaction begins and carbon dioxide evolves. The substances in *solution* come in closer contact and consequently react easily.

In some cases contact between reactants is not the sufficient condition for the reaction to take place ; *heating* is essential. Powdered iron and sulphur may be kept in contact without any chemical reaction taking place. But if the mixture be heated, reaction takes place and ferrous sulphide is formed. Heat splits up mercuric oxide into mercury and oxygen ; the action of heat is sufficient here to cause the reaction to occur.

Coal burns in air. But pulverised coal burns more readily than lump form of it. This is because the surface area of contact between powdered coal and oxygen of the air is greater than that between lump coal and oxygen. An increase in the *surface of contact* leads to an increase in the reaction rate. Thus by bringing the substances into contact a reaction may be initiated and the surface area of contact influences the reaction. That is why the solid reactants are often powdered to have the reaction rate considerably increased.

Application of heat, as has already been stated, is one of the methods of bringing about chemical reactions. Heating generally leads to an increase of temperature. Raising of temperature favours endothermic reaction, while lowering of temperature favours the exothermic one. The synthesis of ammonia from its elements ($N_2+3H_2 \rightleftharpoons 2NH_3$) being an exothermic reaction, rise of temperature decreases the quantities of ammonia. The synthesis of nitric oxide from its elements ($N_2+O_2 \rightleftharpoons 2NO$) on the other hand is an endothermic change. Hence the higher is the temperature, the greater is the possibility of obtaining nitric oxide. It is to be noted, however, that the velocities of practically all reactions increase rapidly with rise of temperature.

Pressure may sometimes brings about chemical changes. Ordinary crackers (Bhui patka) which are made of potassium chlorate and arsenic sulphide or sulphur explodes on percussion. Here we find an example of a chemical change brought about by pressure. Finely powdered lead and sulphur unite under high pressure and lead sulphide is formed : $Pb+S=PbS$.

Increase of pressure favours the change which occurs with diminution of volume ; while decrease of pressure favours that change which occurs with increase in volume. In the synthesis of ammonia from nitrogen and hydrogen diminution of volume occurs, $N_2+3H_2 \rightleftharpoons 2NH_3$. Hence, the higher is the pressure, the greater is the amount of ammonia formed and in actual practice pressure as high as 200 atmospheres is applied. The rate of a reaction is proportional to the concentration (or pressure) of the reactants. As with the passage of time the reactants are used up, their concentration falls, the process becomes slower and slower.

A chemical change may also be brought about by *sound* only. Sound produced by the explosion of mercury fulminate near acetylene gas brings about the decomposition of the latter into carbon and hydrogen.

Many a chemical action is induced by *electricity*. Fused sodium chloride is decomposed by it into sodium and chlorine. Nitrogen and oxygen combine together under the influence of electric arc forming nitric oxide.

In some cases no chemical action takes place in absence of *light*. Hydrogen and chlorine which unite with explosion in direct sunlight fail to react in the dark : $H_2 + Cl_2 = 2HCl$.

Potassium chlorate required to be heated to a very high temperature to yield oxygen produces the same at a comparatively low temperature on being mixed with a little manganese dioxide. Manganese dioxide here accelerates the decomposition of potassium chlorate, itself remaining unchanged in mass and composition. There are many such cases in which a substance by being present in small amount alters the rate of the reaction. But the substance remains unchanged in mass and composition in the process. The phenomenon is called *catalysis* and the substance that alters the rate is called a *catalyst* for that reaction. It should, however, be noted that a catalyst cannot initiate a reaction, it can only influence a reaction already occurring.

We conclude by saying that contact, heat, pressure, sound, electricity, light, etc. are the means of inducing reactions : one or more of which bring about a chemical change. And apart from the nature of the reacting molecules, the factors that influence a reaction are temperature, pressure. radiation, catalyst, concentration of the reactants and their surface area of contact.

1-5. Identification of matter :

A matter is recognised by its characteristic qualities known as properties. At first we can divide matter into three categories namely, solid, liquid and gas. Gold is bright yellow, copper is red, copper sulphate is blue, etc. All are solids ; *colour* of the substance may act as a guide in identifying a matter. The characteristic bright lustre of gold helps us to recognise it. Mercury is a silver-white liquid, both bromine and water are liquid ; but bromine is deep red while water is colourless.

Hydrogen, oxygen, etc. are colourless gases. Chlorine is a greenish yellow gas. No difficulty is felt in recognising iodine vapour by its violet colour. In some cases the *odour* of the substance may help. We cannot distinguish between oxygen, ammonia and hydrogen sulphide simply by observing with the eyes. Odour gives us very useful indication here. Ammonia is characterised by its pungent smell, while rotten egg like smell is indicative of hydrogen sulphide. Oxygen is devoid of any odour. We usually recognise cocoanut oil, kerosene, petrol or spirit, etc., by their characteristic odour. A white solid—whether it is camphor or naphthalene—can be easily perceived through the sense of smell. *Tastes* of diverse nature may help us in the identifiction of matter. One may identify a white solid as sugar, if it tastes sweet. Sweetness is one of the common properties of sugar. Common salt and sugar can easily be distinguished by their characteristic tastes. Both glycerine and water are colourless, odourless liquids. The difference, apart from the slight sweetness of glycerine, may be perceived by the sense of *touch*,—glycerine being a sweet syrupy liquid. Some substances crumble into fine powder under *hammer*. Some substances again can be hammered into thin sheets. Gold, silver, etc. for examples, can be hammered into very thin foils. Magnetic substances, such as iron, cobalt, etc. are attracted by a *magnet*, which, therefore, can indicate the presence of these substances.

The *solubility* of substances is often a useful guide. Common salt, sugar, copper sulphate, etc. are soluble in water. Sulphur is insoluble in water but dissolves readily in carbon-disulphide. While calcium carbonate is insoluble both in water and carbon disulphide but dissolves in carbon dioxide saturated water. Hydrogen gas is practically insoluble in water, but hydrogen chloride, ammonia, etc, are highly soluble. Pure solids and liquids are characterised by their distinct *melting* and *boiling* points respectively. In many a case these help us in recognising the substances.

We have so far discussed how matter differs in physical properties. But it is to be noted here that it often gives us only an idea of the nature of the substance, a complete identification of which demands its specific chemical properties to be investigated. With this object in view the behaviour of substances towards heat and certain specific chemical reagents, such as acids, etc, are observed.

Different substances usually behave differently towards heat. Some undergo simply a physical change and some a radical one. Red mercuric oxide changes into shining mirror of mercury by heat and a colourless, odourless gas is evolved in the process. The evolved gas rekindles a glowing chip of wood. Mercuric oxide on heating decomposes into metallic mercury and oxygen. Smell of rotten eggs evolves, when dilute sulphuric acid comes in contact with small pieces of ferrous sulphide. The obnoxious gas is known as hydrogen sulphide. But a colourless, odourless gas evolves, if metallic iron instead of ferrous sulphide is used. A lighted taper if held in the evolved gas, the latter burns and the taper gets extinguished. The gas that evolves is hydrogen. On heating manganese dioxide with conc. hydrochloric acid a greenish yellow gas, which smells like bleaching powder evolves. The evolved gas is chlorine. But potassium permanganate liberates chlorine from the same acid at normal temperature, no heating is required. On adding silver nitrate to an aqueous solution of common salt a white precipitate appears. The ppt. is silver chloride. Sodium sulphate gives no precipitate under identical conditions. In each case cited above a radical change of the substance or substances concerned has taken place. Take for example the decomposition of mercuric oxide. Mercury, oxygen and mercuric oxide are not one and the same thing. Hence it is a chemical property. The manganese dioxide is black or insoluble in water does not stand as a sufficient proof for its identification. Chemical properties such as reaction with hydrochloric acid, etc. are to be considered. In fact, all the specific properties, both physical and chemical, should be considered as a whole before any conclusion can be drawn.

1-6. Classification of pure substances :

Pure substances are classified inso *elements* and *compounds*. As a result of elaborate investigations it has been determined.

that there are a little over one hundred distinct species or otherwise called the fundamental substances that compose all kinds of matter, familiar or rare. The simple species or fundamental substances are called *elements*. A substance that cannot be decomposed into simpler substances is an elementary substance or *element*. Gold, silver, hydrogen, oxygen, nitrogen, etc. are some of the well known elements.

An electric current decomposes acidulated water into hydrogen and oxygen. Hydrogen and oxygen differ from each other in properties and they bear no resemblance to water. Water is a familiar compound. A substance that can be decomposed into two or more simpler substances is termed a *compound*. It may be defined in another way. A substance formed by union of two or more elements in a definite proportion by weight is termed a *compound*. Common salt, mercuric oxide, carbon monoxide are a few familiar compounds. All substances are aggregates of their molecules. If a hydrogen molecule can be broken down, two indentical atoms *i e.*, hydrogen atoms are formed. A molecule of carbon monoxide on the other hand yields different atoms—an atom of carbon and an atom of oxygen. From what has been said we can infer that an *element* is a substance that contains like atoms only ; while a *compound* is a substance that consists of atoms of different kinds.

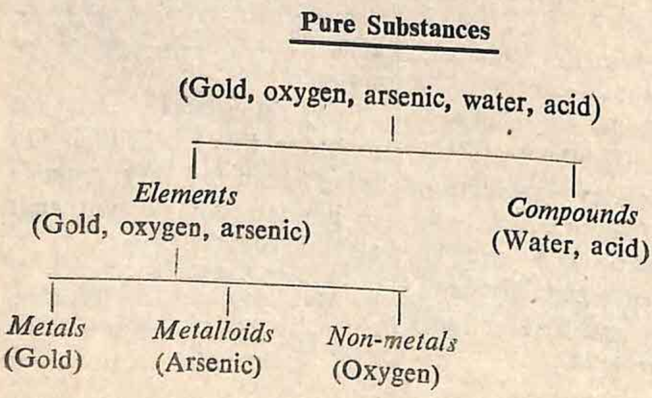
On the basis of properties elements are classified into *metals* and *non-metals*. From time immemorial, man, however, has been putting some substances to his uses. These substances generally seem to be hard but not brittle. They may be hammered into sheets, drawn into wires or melted and cast into moulds. Man was perhaps first acquainted with the uses of gold and copper. The art of using bronze an alloy of copper and tin, was known long before the birth of Christ. The ancient Romans were familiar with the use of lead. Substances like copper, gold, tin, bronze, iron, etc, were, in fact, in the service of man in all countries where old civilization thrived. The special properties for which these substances were considered essential are the characteristic qualities of *metals*. This distinction between metals and non-metals though convenient is an arbitrary one.

Nature seems to abhor abrupt transitions, and shades off her bounding lines, as a few elements arsenic. antimony, etc., for examples, possess some metallic as well as non-metallic properties. These elements are called *metalloids*. The important differences between the two groups are, however, described here in a tabular form :

metals	*Non-metals*
(i) Metals are usually solid and possess lustre known as metallic lustre. Gold, copper, silver, etc. are solid & bright. Exception : Mercury, though a metal, is a liquid.	(i) Non-metals are generally liquid or gaseous and have no lustre ; e.g., bromine is a liquid ; hydrogen, oxygen, nitrogen, etc. are colourless gases. Exception : Iodine, phosphorus, sulphur, etc. are solid and iodine possesses a lustre like metals.
(ii) Metals are good conductors of heat and electricity. Silver is the best in this respect, while copper stands second. You perhaps know that wires of copper (or aluminium) are used for carrying electricity.	(ii) Non-metals are usually poor conductors of heat and electricity. A piece of charcoal can be heated to redness by holding in between fingers, while a coin made of metals cannot be held like this in a flame for such period. Exception : Graphite, though a non-metal, is a good conductor of heat and electricity.
(iii) Metals are comparatively heavy *e.g.*, lead, mercury, etc. Exception : Sodium, potassium, etc. are lighter than water.	(iii) Solid non-metals are usually light. Exception : Iodine, though a non-metal is heavier than some of the metals.
(iv) Metals are malleable and ductile ; can be made into sheets under hammer and drawn into wires. When struck with a hammer it emits a peculiar sound known as metallic sound. Exception : Metals like antimony and bismuth are brittle.	(iv) Non-metals are not sonorous, they are brittle. On being hammered they crumble to powder and do not emit metallic sound.

Metals	*Non-matals*
(v) Metals have high melting points.	(v) Non:metals have low melting-points. Exception: Non-metals carbon and silicon have high melting points.
(vi) Metals are electro-positive. In electrolysis metals get deposited at the negative electrode, *i.e.*, cathode.	(vi) Non-metals are electro-negative. Non-metals liberate at the positive electrode, *i.e.*, anode. Exception: Hydrogen though a non-metal liberates at the cathode.
(vii) Metallic oxides are usually basic in nature and form salts on reaction with acids.	(vii) Oxides of non-metals are usually acidic and never basic in nature, and if soluble form acids on reaction with water.
(viii) Some metals displace hydrogen from acids and form salts.	(viii) Non-metals usually have little affinity for acids but if acted upon by acids, they do not yield salts.
(ix) With hydrogen *a few* metals form compounds known as hydrides, which are unstable, *e.g.*, NaH, CaH_2 etc.	(ix) Non-metals form hydrides, which are generally stable, *e.g.*, NH_3, H_2O etc.

From the discussion just finished, the classification of pure substances is shown in the following form :



Questions

1. What is matter? What are the different states of it? Give suitable examples.

2. What are the characteristics of physical and chemical changes? Give your answer with suitable illustrations.

3. Show by some examples how matter is identified by its physical and chemical properties.

4. Compare metals with non-metals.

5. Write in short what you know about the factors that induce and influence a chemical change.

6. Classify giving reasons the following changes as physical or chemical:

(i) A brown crust is formed on the surface of a piece of iron when left to itself in a moist place; (ii) a platinum wire, when heated become white hot and emits light; (iii) water, if sprayed on a lump of quicklime the latter crumbles to fine powder; (iv) cooling freezes water into ice and if the ice is alowed to melt, water is obtained again; and (v) passages of electric current through common salt leading to formation of sodium and chlorine.

7. Indicating the differences define and illustrate:

(i) elements and compounds. (ii) endothermic and exothermic reaction.

Sample questions for informal objective type tests

A. Alternate response type

(i) *True or False type*:

(a) A catalyst cannot initiate a reaction, it only influences the rate of it. —

(b) Non-metallic oxides are usually basic in nature. —

(c) A chemical change is always attended either with the evolution or absorption of heat. —

(*ii*) *Yes or No type*:

(a) Has the gaseous substanc a shape and volume of its own? —

(b) Is there any element that possesses some of the properties of both metals and non-metals? —

B. Recall Type:

(*i*) The substance that cannot be decomposed into simpler substances is called an – substance. —

(*ii*) The change that does not involve any change in composition is called a — one. —

(*iii*) The substances in solution come in — contact and consequently react sometimes easily. —

C. Multiple choice type :

Selection type : [Pick out a suitable word from column II and fill up the gap in column I]

Column I	Column II
(a) A liquid possesses a definite — and mass.	Exothermic, elements, volume, arbitrary, compounds.
(b) Water, common salt etc. are some of the well known —,	
(c) Oxygen, nitrogen, hydrogen, etc. are examples of familiar —,	
(d) The classification of elements into metals and non-metals is an — one.	
(e) The reaction attended with the evolution of heat is called an — reaction.	

CHAPTER TWO

Solution and solubility

An interesting property of water is its ability to dissolve substances. Ocean is an aqueous solution, thousands of components such as ions of metals and non-metals, complex inorganic ions and different organic substances. etc. are present in it, It was in this solution that the first living beings came into sight and this provided them with the necessary ions and molecules for their life and growth. The part played by solutions are, in fact, important for both life and industry. Suffice it to mention that the process of assimilation of food involves dissolution of substances.

2-1. **Solution :** Shake thoroughly a little sugar with water, the former disappears. But the sweet taste of water indicates the presence of sugar in it. Sugar is said to be dissolved in water and a *solution* has been formed. If instead of sugar a little sand be taken, a portion of it remains suspended in water and a portion gets deposited at the bottom, We say that the sugar is soluble and the sand is insoluble in water. But it is to be noted that there is no substance which is absolutely insoluble. Sand even has some solubility, however, small it may be. The solubility of sand in water is so small that we call it insoluble. If 100 gm. water dissolves less that 0·01 gm, of a substance, we generally call it insoluble.

Take the case of sugar-solution in which sugar and water formed a uniform or otherwise called *homogeneous* mixture. The proportion of sugar and water in any or every part of the mixture is one and the same. The substance that gets dissolved is usually called the *solute* and the dissolving medium the *solvent*. In the example under consideration, water is the solvent and sugar is the solute.

∴ Solute+solvent=Solution

A *solution* is, therefore, a homogeneous mixture of two or more substances.

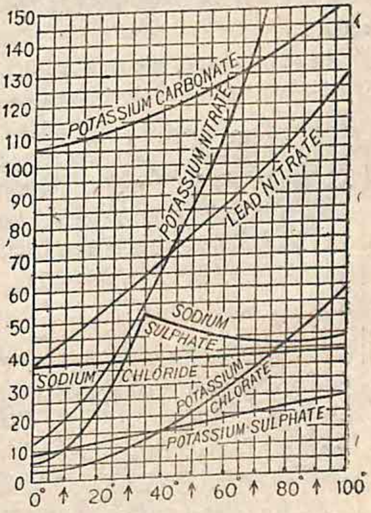
Unless otherwise stated by the term 'solution' an aqueous solution is generally meant. But water is not the only solvent. A substance insoluble in water may dissolve in another liquid. Sulphur is insoluble in water but dissolves readily in carbon disulphide. A solution may be solid or gaseous. Air, for example, is a homogeneous mixture of nitrogen and oxygen mainly ; and hence it may be called a *gaseous solution*. Again ethyl alcohol is a liquid and it is miscible with water in all proportions. Which is the solvent here ?—Water or alcohol ? The use of the terms, in fact, is a mere matter of convenience and there is no real difference between a solute and a solvent. In case of a solution the component which is present in larger amount than the other may be called the *solvent* and the other the *solute*. If, however, one component of the solution is a liquid, it is customary to regard the liquid as the solvent. The discussion here has been primarily confined to these solutions and water is the most generally used liquid.

To a definite amount of water at room temperature go on adding sugar, a little at a time, with constant stirring, the sugar at first dissolves readily. But after some time a point will be reached when no more sugar dissolves but simply settles down unchanged. It is evident that the solution at the particular (room) temperature contains the maximum possible amount of the dissolved substance. The solution is said to be saturated with sugar at the temperature of the experiment. On adding more water or on heating, more sugar dissolves ; the amount of solute dissolving in a given amount of solvent increases with the rise of temperature. On cooling the solution, it is generally found that definite quantity of solute separates out with the fall of temperature. To attribute a quantitative meaning to solubility, the amount of solvent is to be *fixed* up and the temperature stated. This fixed amount is *100 gm*. of solvent. We will discuss it a little later.

At a given temperature the ability of a definite quantity of a solvent to dissolve a substance is fixed. A solution is said to be *saturated* at a given temperature, if the composition of the solution does not change in presence of solid solute. A saturated solution, therefore, contains maximum possible amount of the solute in the dissolved state and this solution will be regarded as saturated so long as the temperature remains unchanged. If the solution, contains less than what is required to form a saturated solution, it is said to be *unsaturated.* But if it somehow contains more solute, it is said to be *supersaturated.* The concentration of an unsaturated solution, therefore, increases when some solute is added to it. The supersaturated solution is unstable in nature ; addition of solid solute results in decrease in concentration (solute being thrown out of solution). Stirring the solution or scratching the wall of the container or the presence of a little dust causes the excess solute to be deposited. The composition of a compound, it is to be noted, is fixed, while the composition of a solution may vary.

2-2. **Solubility :** There is a limit to the maximum amount of solute that can be dissolved in a definite quantity of solvent at a particular temperature. The measure of solubility of a substance under a given condition is the concentration of its saturated solution. The *solubility* of a solute in a solvent at a given temperature is usually defined as the number of grams of the solute required to saturate the *fixed amount i.e.*, 100 gms. of the solvent at the temperature. If, for example. m gms. of solute can saturate y gms. of solvent, at t°C, then its solubility in the given solvent at $t^\circ\text{C}=\frac{100m}{y}$.

The solubility usually varies with temperature and this variation is shown by means of a curve or graph in which abscissae represent temperatures and ordinates the solubilities. Such a curve as this which represents solubilities at various temperatures (the relation of solubility of a substance to temperature) is called the *solubility curve* of that substance. Solubility curves of some familiar salts are shown in the adjacent figure.



Solubility curves

The solubility of a substance at a given temperature can be read off directly from the curve. Secondly, it enables us to have a comparative idea of the solubilities of different substances. Thirdly, the amount of solute that will separate out on cooling the saturated solution (at a particular temperature) of it to a lower one can be calculated. Fourthly, when a solution contains more than one solute, the curve tells us which of the solute will separate out first when subjected to evaporation.

Questions

1. What do you mean by the terms : 'solvent', 'solute' and 'solution' ? Explain with examples.

2. Write in brief what you know about the terms : saturated—, unsaturated— and super-saturated solutions. How does a solution differ from a compound ?

3. What is meant by solubility ? Is it essential to mention the temperature while stating the solubility of a substance ?

4. Define 'solubility'. What is meant by a solubility curve ? State its usefulness.

Sample Questions for Objective type tests

1. Alternate response type

True or False type :

(i) A homogeneous mixture of two or more substances is called a solution. —

(ii) The composition of a solution is invariable —

(iii) Air may be called a gaseous solution. —

2. Completion type

A saturated solution contains — (a) possible amount of the — (b) in the — (c) state and this solution will be regarded as satuarted so long as the — (d) remains unchanged.

—(a)
—(b), —(c)
— (d)

3. Recall type

(i) There is no — difference between a solute and a solvent. —

(ii) A solubility curve represents the relation of solubility of a substance to —. —

CHAPTER THREE

Symbols, Formulae, Valency & Equations

[In balancing a chemical equation the method of writing chemical formulae and consequently the idea of valency are essential. These connected topics have, therefore, been dealt with here along with the subject matter included in the syllabus.]

3-1. Symbols and formulae :

The modern simple system of representing the full mame of an element in brief is due to Berzelius, a Swedish chemist. In many cases the initial letter (capital) of the common name of the element is used as abbreviation for it. For example H stands for hydrogen, O for oxygen, N for nitrogen, etc. In case of two or more elements having the same initial letters, two letters are used, one of the elements, however, is represented by initial letter only. Thus, Boron—B, Barium—Ba, Bromine—Br etc.

These are referred to as *symbols*. The symbol is thus an abbreviation for the full name of an element.

Any kind of molecule, elementary or compound is now represented by a combination of symbols and numerals. This is referred to as *formula* of the substance. A molecule of an element consists of one kind of atom (one or more), while the molecule of a compound always consists of different kinds of atoms. The number of atoms per molecule of the element is known as the *atomicity* of the molecule. If the molecule of an element contains one atom, then the molecule is represented by the symbol only, *i.e.*, in case of mono-atomic molecule the symbol represents also the formula. The formulae of carbon and mercury which are mono-atomic are C and Hg respectively. In case of poly-atomic molecule the number of atoms is put as subscript. Each of the molecules of nitrogen, hydrogen, oxygen, etc. consists of two atoms ; and their formulae are represented by N_2, H_2 and O_2 respectively. To build up the formula of a compound the symbols of the constituent elements are written side by side and the number of atoms of each is indicated by putting numerals to the lower right of the symbols. In a molecule of water there are two atoms of hydrogen and one atom of oxygen. H_2O, therefore, denotes the formula of water. NH_3 stands for ammonia, sulphuric acid has the formula H_2SO_4. The symbol itself, it is to be noted, represents one atom of the element.

A molecule of hydrogen, as has already been stated, consists of two atoms of hydrogen and its formula is written as H_2. But 2H represents two atoms of hydrogen. By $3H_2$ we mean three molecules of hydrogen and it contains six atoms of hydrogen, but 6 atoms of hydrogen are expressed as 6H. $10H_2O$ means ten molecules of water. Similarly $2HNO_3$ would mean two molecules of nitric acid. Hence, to denote the number of atoms and molecules, the required numeral is placed before (*i.e.,* left-hand side of) the symbol and formula.

3-2. Valency : The combining capacity of an atom or radical is by and large fixed. This capacity is what is known as *valency*. The idea should be made more clear as its relation to chemical formulae must be thoroughly understood. The valency of an element is the combining capacity of an atom of the element and is measured by the number of atoms of hydrogen it will combine with or displace. Nitrogen, for example, combines with hydrogen forming ammonia (NH_3) ; one atom of nitrogen here combines with three atoms of hydrogen and hence valency of nitrogen is 3. Again an atom of zinc displaces two atoms of hydrogen from hydrochloric acid forming zinc chloride ($ZnCl_2$), the valency of the element zinc by definition is, therefore, 2. Valency of an element that neither combines directly with hydrogen nor displaces it from its combination is determined with reference to other elements of known valency. Gold, for example, neither combines with hydrogen nor displaces it from its combination. But it combines with chlorine forming auric chloride ($AuCl_3$). Valency of chlorine with reference to hydrogen is 1 and hence the valency of the element gold, an atom of which combines with three atoms of chlorine forming auric chloride must be 3. Most elements can combine with oxygen and hence it becomes often convenient to find out valency by the study of oxides. Oxygen, we know, is bivalent with reference to hydrogen the original standard. Twice the number of oxygen atoms that can combine with an atom of the element in question is its valency. An atom of magnesium combines with one atom of oxygen forming magnesium oxide (MgO) and hence magnesium is bivalent. Argon, helium, etc., the so-called inert gases, have no combining capacity and hence they are regarded as zero (0) valent elements. The highest valency known being 8, the valencies range between 0 and 8, and hence, in all nine valencies are there. Valency, it is to be noted, is always a whole number. Valency may vary, some elements exhibit different valencies *i.e.,* they have variable valency. Copper is univalent in cuprous oxide (Cu_2O) but bivalent in cupric oxide (CuO).

Groups of atoms such as NO_3, SO_4, PO_4 etc., that enter into the composition of chemical compounds and participate in chemical reactions as an integral whole are known as *groups* or *radicals.* The concept of valency already described is equally applicable to the radicals also. In nitric acid (HNO_3), nitrate radical (NO_3) is mono-valent. since it is combined with one atom of hydrogen. Similarly sulphate group (SO_4) is bivalent in sulphuric acid (H_2SO_4). Valencies of some important elements and radicals have been shown in tabular form in the appendix.

Writing formula : The knowledge of the valencies of elements is useful in arriving at the formulae of compounds.

When an element A combines with another element B to form a compound, then the number of atoms of A and B in the resulting compound will be in inverse ratio (usually in its lowest terms) of their valencies. The same applies to groups also.

Applications : (i) Write down the symbols of the constituent elements side by side (metals or more electropositive elements or radicals are to be indicated first).

(ii) Indicate the valencies of each by putting numerals at the top and a little towards right of the symbol.

(iii) Transfer the numerals cross-wise. In this case also '1' is not to be indicated.

The following examples will make the point more clear :

Names of compounds.	Symbols with valencies	Formulae
Calcium chloride	Ca^2Cl^1	$CaCl_2$
Ferric oxide	Fe^3O^2	Fe_2O_3
Calcium phosphate	$Ca^2PO_4^3$.	$Ca_3(PO_4)_2$
Potassium sulphate	$K^1SO_4^2$	K_2SO_4

If, however, the numerals representing valencies be divisible by a common factor, then they are reduced by dividing with such factor and the resulting numerals are then transposed as before. Below is given a list of compounds, which is illustrative of this point.

Names of compounds	Symbols with valency	Reduced numerals	Formulae
Calcium sulphate	$Ca^2SO_4{}^2$	$Ca^1SO_4{}^1$	$CaSO_4$
Carbon dioxide	C^4O^2	C^2O^1	CO_2
Magnesium carbonate	$Mg^2CO_3{}^2$	$Mg^1CO_3{}^1$	$MgCO_3$

3-3. Chemical equations :

In a chemical change a new substance or substances are always formed. The change may conveniently be represented by a *chemical equation*. The substance or substances undergoing the change and the new substance or substances formed thereof can be represented by the respective symbols and formulae. The substances undergoing the change or the reacting partners are indicated on the left-hand side of a sign of equality (=) and the new substance or substances, *i.e.*, the products on the right of it. The reactants and so also the products are connected by plus (+) signs. The number of atoms of each element on either side of the sign of equality are made one and the same. The plus sign on the left means 'reacts with' while the same on the right of the sign of equality denotes 'and'. The sign of equality itself means 'to produce.' This is how a chemical change is shown by a chemical equation. Hence, a sort of *short-hand* representation of a chemical change in brief is known as a chemical equation. Thus, the reaction between hydrogen and chlorine to produce hydrogen chloride is represented by the equation :

$$H_2+Cl_2=2HCl$$

Significance : It has both qualitative and quantitative significance.

(*i*) It tells us the nature of the reactants and products. The equation $H_2+Cl_2=2HCl$, for example, indicates that hydrogen reacts with chlorine to produce hydrogen chloride.

(*ii*) It also denotes the relative amounts by weight of all the substances involved in the change. In the above reaction

1 molecule of hydrogen combines with 1 molecule of chlorine to produce 2 molecules of hydrogen chloride.

Hence, we can say from the above equation that 2 parts by weight of hydrogen combine with $(2 \times 35{\cdot}5)$ parts by weight of chlorine to produce $\{2 \times (1+35{\cdot}5)\}$ parts by weight of hydrogen chloride.

(*iii*) If any of the reactants or products be gaseous, the volumes of such gases can also be known from the equation. Again, if the reactants and products are all gaseous, the equation indicates the volumetric relations existing among them. In the example under consideration the equation tells us that (under identical condition of temperature and pressure) 1 volume of hydrogen combines with 1 volume of chlorine to produce 2 volumes of hydrogen chloride.

To write a correct chemical equation : A chemical equation expresses a chemical change or reaction that does really occur, mere balancing does not make it correct. Hence to write an equation the names of all the substances involved in the change must be known. Since the law of conservation of mass holds good, balancing of the equation is essential. All substances should be expressed in their molecular forms. For mono-atomic elements like sodium, mercury, etc. and for elements (solid elements in the free state *e.g.*, Zn, Fe, etc.) in which the number of atoms per molecule is sometimes, though not always indefinite or variable the respective symbols should stand for their formulae.

Applications : (*i*) Write down the names of the reactants and products. Connect the reactants and so also the products by plus (+) sign. Separate them by putting an arrow (→) so that the reactants lie on its left. (*ii*) Represent the elements by symbols and the compounds by formulae. (*iii*) Now balance so that there are exactly the same number of atoms of each element on the left as on the right side of the arrow sign. (*iv*) Lastly express each substance in the molecular form and replace the arrow by a sign of equality.

Illustrations : (*a*) Hydrogen and oxygen combine forming water.

$$\text{Hydrogen} + \text{Oxygen} \rightarrow \text{water} \quad \cdots(\text{i})$$

$$H + O \rightarrow H_2O \quad \cdots(\text{ii})$$

$$2H + O \rightarrow H_2O \quad \cdots(\text{iii})$$

$$2H_2 + O_2 = 2H_2O \quad \cdots(\text{iv})$$

(*b*) Potassium chlorate decomposes on heating, potassium chloride is formed and oxygen is liberated.

$$
\begin{array}{lcccl}
\text{Potassium Chlorate} & \rightarrow & \text{Potassium chloride} & + & \text{Oxygen} \quad \text{(i)} \\
KClO_3 & \rightarrow & KCl & + & O \quad \cdot\cdot\ \text{(ii)} \\
KClO_3 & \rightarrow & KCl & + & 3O \quad \cdots \text{(iii)} \\
2KClO_3 & = & 2KCl & + & 3O_2 \quad \cdot\cdot\ \text{(iv)}
\end{array}
$$

We summarise by saying that equation be so adjusted (if not automatically balanced) by placing the smallest possible suitable numerals before the corret formulae (sometimes symbols) of the substances, as the number of atoms of each element are exactly the same on both sides of the sign of equality ; and in the case of simple equation it is convenient to achieve the same objective by the method stated above.

Questions

1. Define and illustrate : symbol, formula and valency.

2. What is meant by a 'formula' ?

Write down the formulae of the following compounds :

cupric sulphate, cuprous chloride, mercuric sulphate, magnesium hydroxide, potassium hydroxide, calcium carbonate, calcium bi-carbonate, aluminium phosphate, calcium phosphate, potassium phosphate, aluminium sulphate, mercurous chloride and ferric oxide.

3. What is meant by a chemical equation ? State its significance. Why should you balance an equation ?

4. Illustrate with an example the method of writing a chemical equation.

5. State fully what informations are afforded by the equation : $N_2 + 3H_2 = 2NH_3$.

6. Balance the following :

(i) $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$, (ii) $Fe_2O_3 + Al \rightarrow Fe + Al_2O_3$, (iii) $MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$, (iv) $Ca_3(PO_4)_2 + SiO_2 \rightarrow CaSiO_3 + P_2O_5$.

7. Frame equations for :

(i) magnesium + sulphuric acid = Magnesium sulphate + hydrogen,

(ii) iron + steam = ferroso-ferric oxide + hydrogen,

(iii) magnesium + nitrogen = magnesium nitride,

(iv) potassium nitrate = potassium nitrite + oxygen.

Sample questions for objective type tests

A. Alternate response type

True or False type

(a) Mere balancing does not make an equation correct. —

(b) Valency is a constant property. —

(c) The number of atoms of each elemant on both sides of an equation must balance. —

B. Recall type

(i) – is the abbreviation for the full name of an element. —

(ii) Molecule of a compound always consists of atoms of — element, —

C. Multiple choice type

[Pick out a sutable word from column II and fill up the gap in column I]

I	II
(i) Oxygen is a – element.	actually, divalent, and, formula.
(ii) Symbol Hg stands also for the — of mercury.	
(iii) Correct chemical equation represents the change that does — occur.	
(iv) The plus (+) sign on the right hand-side of the sign of equality denotes —.	

CHAPTER FOUR

Electrolysis

4-1. Electrolytes and Electrolysis :

You all perhaps know that all substances are not capable of conducting an electric current. Mica, rubber, wood, etc., are some of the examples of it ; while gold, silver, copper, etc., are good conductors of electricity, they allow current to pass through. Hence, on the basis of their behaviour towards electric current substances are broadly classified as *conductor* and *non-conductor*. The conductors again are of two types.

Some while carrying current do not undergo any permanent or otherwise called chemical change, One of the most familiar instances of this type of conductor is the filament of an electric lamp. The metallic wire of which the filament is composed gradually gets heated due to the passage of current and becomes white hot and emits light. When the supply is cut off, the original state is restored. These are termed as *metallic conductors* though a few (*e.g.*, gas carbon) of them are not definitely metals.

Then again some, which either in a state of solution or in a state of fusion conduct electric current and during the process of conducting electric current undergo a chemical change are referred to as *electrolytes*. Aqueous solutions of many salts, strong acids and alkalies are good electrolytes. A substance which conducts electricity with simultaneous chemical change is, therefore, called an electrolyte.

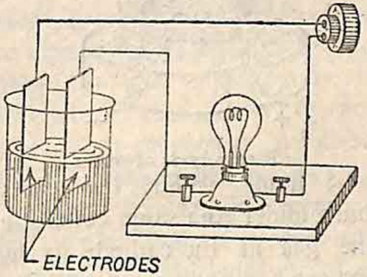
Hence, substances such as metals, metallic alloys, gas carbon, etc., are conductors but not electrolytes, since they do not undergo decomposition while carrying current. But substances like sugar, glycerine and a host of other organic compounds neither belong to this group nor are electrolytes under any circumstances and these are called *non-electrolytes*.

The process of decomposition of a substance in a state of solution or fusion by an electric current is called *electrolysis*.

The vessel in which electrolysis is carried out is called an *electrolytic cell* or *voltameter* and the conductors—rods, wires or sheets—by which the current enters or leaves the electrolytic bath are called *electrodes*. The electrode connected with the positive pole of the source is known as *anode* and that with the negative is *cathode*.

The different behaviour of substances towards electric current may easily be shown by the experiment described below.

Two electrodes of graphite or of some metal (nickel or platinum preferable) are connected with the source of current as shown in the accompanying figure. The free ends of the electrodes are dipped into dry and pure common salt in the beaker and the current is switched on. The lamp does not light up. The same result is obtained, if the sodium chloride is replaced by very pure water. But if a little of the above salt or caustic soda be dissolved in this water or acids like hydrochloric or sulphuric be mixed with it and the current is switched on, the lamp will light up and burn more or less brightly. Dry and pure common salt, very pure water are, therefore, not capable of conducting current. But the wire, electrodes *i.e.*, the materials of which they are made and the aqueous solutions of salts and of acids and bases conduct current. Besides, by the same experiment again a rough estimate about the electrical conductivity of different solutions may be made from the brightness of the lamp.

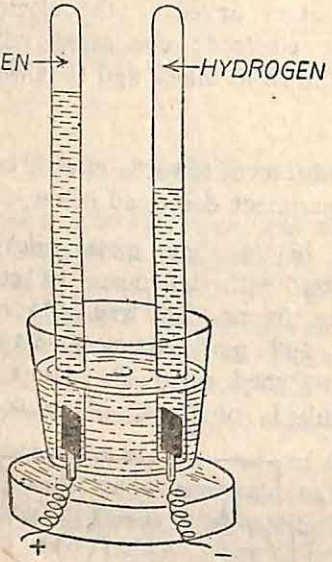


Different behaviour of substances towards electric current

The aqueous solution of salts and of acids and bases undergo chemical changes while carrying current. On passing current through water acidulated with sulphuric acid by means of platinum electrodes water is decomposed. Hydrogen is liberated at the cathode and oxygen at the anode. This phenomenon is what is known as *electrolysis of water*.

$$2H_2O \rightarrow 2H_2 + O_2$$

Experiment : Two platinum foils fused at the bottom of a glass basin stand vertically. The foils are kept completely submerged under water by pouring pure water in the basin In order to make this water conducting a little sulphuric acid is mixed with it. Two tubes filled with this acidulated water are kept inverted over the foils in the basin. The foils are connected to the positive and negative poles of a battery and the current is switched on. The gas collects in the tubes by the displacement of water. A glowing chip of wood if introduced into the gas collected at the cathode extinguishes and the gas burns with a pale blue flame. While the gas collected at the anode does not burn under the same condition but rekindles the chip. Hence, the gas at the cathode is hydrogen and that at the anode is oxygen.



Electrolysis of water

Ionisation : Pure water for all practical purposes is a non-conductor of electricity. Liquid hydrogen chloride, pure and dry common salt are not conductors, nor also is the dry caustic soda. But an aqueous solution of each of these substances can conduct current. Evidently some kind of change must have taken place on dissolution of the substances in water, as a result of which they become capable of conducting electric current.

To account for this, Arrhenius, the Swedish scientist, put forward a theory known as ‘theory of electrolytic dissociation’ in 1887. In water or other ionising solvent the electrolytes break up spontaneously to a greater or smaller degree into two or more charged particles. These charged particles are called ions. The process is reversible, the resulting ions and the undissociated molecules remain in equilibrium. Each substance gives two kinds of ions—positive and negative. This spontaneous reversible dissociation is known as *electrolytic dissociation* or simply *ionisation*. Electricity is carried by charged particles *i.e.*, free ions and not by undissociated molecules. Anhydrous hydrogen chloride consists of HCl molecules, but it dissociates into hydrogen ions and chloride ions, when dissolved in water. The presence of free ions would account for its ability to conduct electric current. This ionisation is expressed as follows :—

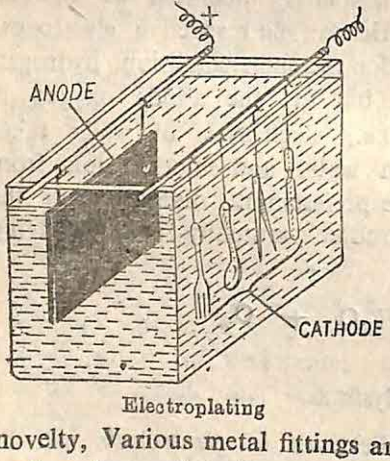
$HCl \rightleftharpoons H^+ + Cl^-$, the process being reversible. the sign of reversibility ($\rightleftharpoons$) has been used instead of the sign of equality. Hence the break-up of electrolytes into ions is not effected by the current. The only effect of the applied electrical potential is to direct the ions to the electrodes of opposite sign.

When direct current is passed through conc, hydrochloric acid, hydrogen ions and chloride ions move to the cathode and anode respectively, They are relieved of their charge, converted into neutral atoms and are ultimately liberated as gaseous (molecular) hydrogen and chlorine at the respective electrodes. With dil. hydrochloric acid under the same condition hydrogen is liberated at the cathode, but at the anode oxygen is liberated instead of chlorine *i.e.*, electrolysis of water takes place. Of the OH^- ion from water and chloride ion from hydrochloric acid the former is preferentially discharged at the anode. The neutral hydroxyl radicals mutually react evolving oxygen :

$$4(OH) \rightarrow 2H_2O + O_2$$

4-2. Application of electrolysis :

Some of the important applications of electrolysis will be described here in brief. If an electric current is passed through a copper sulphate solution acidified with a little sulphuric acid by means of two copper plates connected with the poles of a battery, copper dissolves out from the anode (the plate connected with the positive pole) and gets deposited on the cathode (the plate connected with the negative pole). It thus provides us with a way of putting a coherent, handsome and stable layer of expensive metals like gold, silver, nickel, copper, etc. on an object under appropriate conditions. The process is known as *electroplating*. The object to be plated is first cleansed with caustic soda solution in order to remove fatty and oily matters from its surface. It is then treated with dil-hydrochloric or sulphuric acid to remove the oxide film and washed thoroughly with water. The object thus cleansed is then polished. A suitable salt of the plating metal under appropriate condition is made the electrolyte. The cleansed and polished article is made the cathode and the plating metal the anode. To plate an article with silver, for example, the cleansed and polished article is used as cathode and the plate of pure silver as anode, The two are dipped into a solution of potassium argentocyanide, which here acts as the electrolyte. When the current is switched on, silver dissolves from the anode and gets deposited on the article acting as the cathode. A beautiful coating of silver soon spreads over the surface of the article. To plate an article with gold potassium aurocyanide is used as electrolyte. Gold dissolves from the anode and deposits on the cathode. Nickel plating is done in much the same way as copper. Aim of plating is to prevent corrosion or to add to the beauty of the article or sometimes to achieve both. Nickel film is easily attacked with acids and in the air of towns and cities it soon tarnishes. That is why it has been largely replaced by chromium and chromium plating is a comparative novelty, Various metal fittings are plated with chromium. You perhaps know that silver or gold plated ornaments are in use. Table wares such as knives, forks and spoon, etc. are sometimes coated even with silver or gold.



Electroplating

A process similar to electroplating is known as *electrotyping*. It is the art of preparing electrotypes and reproducing statues or other works of arts. A wax impression or plaster cast of the type, statue, etc. is first made. The surface of it is then made a conductor of electricity by covering it with graphite powder. Copper may now be deposited on the casts used as cathode by electrolysis of copper sulphate solution under appropriate conditions as stated before. This being done it is taken out of the bath and the wax melted off or the shell stripped off. The surface of the deposited layer facing the cast or mould is an exact copy of the object.

Questions

1. Define and illustrate :

Electrolytes and non-electrolytes, electrolysis, anode and cathode.

2. What is meant by 'electrolysis'? Describe the electrolytic decomposition of water.

3. Devise an experiment to demonstrate the different behaviour of substances towards electric current.

Sample questions for objective type tests

A. Alternate response type

(i) *True or False type*

(a) Pure water is a good conductor of electricity. —

(b) Aqueous solution of pure common salt allows the current to pass through. —

(ii) *Yes or No type* :

(a) Does an electrolyte undergo decomposition while carrying an electric current? —

(b) Should the plating metal be used as cathode in electroplating? —

B. Recall type

(i) The vessel in which electrolysis is carried out is called a —. —

(ii) The anode is the positive electrode by which the current — an electrolyte. —

C. Completion type

To plate an article with silver the cleansed and polished article is used as—(a) and the plate of pure silver as—(b) The two are dipped into a solution of — (c), which here acts as the — (d), When the current is switched on, silver deposits on the article used as — (e).

—(a)
—(b)
—(c)
—(d)
—(e)

CHAPTER FIVE

Acids, Bases and Salts

Each of the terms acids, bases and salts signifies a distinct class of compounds and is, therefore, used in a generic sense. They are so interrelated that the properties of any of these cannot be properly dealt with without any reference to the others.

5-1. Acids :

The term is derived from the Latin word 'acidus' meaning sour. Hydrochloric, sulphuric. nitric, etc. are some of the well known acids. The sour taste of various green fruits is due to the presence of acids with some of which we are familiar in our everyday life. Citric acid, for example, is present in all citrus fruits such as lemon. Tamarind contains tartaric acid, etc. etc The question naturally arises why all the substances mentioned here are included under the common-name acid. The answer to this will be found in the properties which are by and large commonly shared by all of them.

(*i*) Most acids more or less taste sour. (*ii*) They generally dissolve in water and the aqueous solutions conduct electricity. (*iii*) They turn blue litmus red.

Expt. Take a test tube with a little water in it. Add a small amount of hydrochloric acid, shake thoroughly and dip a blue litmus paper in the resulting mixture ;—the litmus turns red. Repeat the experiment with sulphuric acid, vinegar, citric acid, etc. The same results are obtained.

(*iv*) Acids consist of two parts : one part contains one or more hydrogen atoms and the other part one or more nonmetallic atoms or groups. It is this second part which attributes distinct qualities to different acids and hence it is referred to as acid radical.

(*v*) Hydrogen atoms of acids may be partially or wholly replaced by certain metals or groups acting like a metal. As a result, the metals or groups unite with the acid radicals and new compounds known as salts are formed.

Expt. See the laboratory method of preparation o hydrogen. Hydrogen atoms of sulphuric acid have been replaced by zinc. The residual solution in the woulf's bottle contains zinc sulphate, which is a salt of zinc.

(*vi*) Acids dissolve certain metals forming salts with them.

Expt. Take a few pieces of copper turnings in a test tube, pour 1 : 1 nitric acid on it and finally apply heat to dryness.

Red copper disappears and a blue transparent solution results. At last a blue coloured residue is left behind. This residue is entirely different from red copper turnings. Its name is copper nitrate and it is a salt of copper. Hence, copper dissolves in nitric acid and a salt named copper nitrate is formed.

(*vii*) An acid reacts with a base forming salt and water. This is the most important property of an acid.

The substances possessing the above properties are described as acids. To form an elementary idea, an acid may be defined as *a compound containing hydrogen, a part or whole of which may be replaced, directly or indirectly, by a metal or a group of atoms acting like a metal to form a salt.*

In the reaction between zinc and dil. H_2SO_4 the former displaces directly and wholly hydrogen of the latter forming a salt named zinc sulphate, $Zn + H_2SO_4 = ZnSO_4 + H_2$. But copper cannot directly replace the hydrogen, the replacement may, however, be effected by an indirect method such as the action of acid upon cupric oxide : $CuO + H_2SO_4 = CuSO_4 + H_2O$. Though this replacement is an indirect one, the net effect is the same in the sense that copper sulphate may be considered to have been formed by replacing the hydrogen of acid by copper, when we think of the relation between copper sulphate and sulphuric acid. Take the case of two compounds ammonium hydrogen sulphate NH_4HSO_4 and ammonium sulphate $(NH_4)_2SO_4$ for consideration. In the first case the hydrogen of sulphuric acid is partly and in the second wholly replaced by ammonium radical.

5-2. Bases :

Compounds possessing the opposite qualities of acids and react with the latter forming salts and water are termed as *bases.* Sodium oxide (Na_2O), calcium oxide (CaO), sodium hydroxide ($NaOH$) are some of the familiar bases. A base also consists of two parts. Generally one or more atoms of metal constitute the first part and it is the part that determines the distinct qualities of a particular base. This portion of a base is called a basic radical. The remaining part contains oxygen atom or hydroxyl (OH) groups.

The most important property of this class of compound is the ability to react with acid forming salt and water. This reaction is known as *neutralisation*. The properties of each are lost, they mutually neutralise each other.

$$\textit{Base} + \textit{Acid} = \textit{Salt} + \textit{Water}$$

$$\underset{\text{Calcium oxide}}{\mathrm{CaO}} + \underset{\text{Hydrochloric acid}}{2\mathrm{HCl}} = \underset{\text{Calcium chloride}}{\mathrm{CaCl_2}} + \underset{\text{Water}}{\mathrm{H_2O}}$$

It is evident from the above that oxides and hydroxides of metals possess this property. Though ammonia and some other compounds are not oxides or hydroxides of metal and form no water but salts only on reaction with acids are also regarded as bases. Water soluble hydroxides are called *alkalis*. Sodium hydroxide, potassium hydroxide, ammonium hydroxide (NH_4OH) are soluble in water and are well known alkalis. All alkalis are, therefore, bases, but all bases are not alkalis. The generic properties of alkalis are : (i) They possess an astringent taste. (ii) The aqueous solution of these substances exhibits a 'soapy touch' and turns red litmus blue.

Expt. Take some caustic soda in a test tube. Shake well with a little amount of water and test the resulting mixture with a red litmus paper. The litmus turns blue.

Hence, an acid turns a blue litmus red and an alkali a red litmus blue. (iii) Aqueous solutions of these conduct electricity. (iv) Being a base in the wider sense an alkali shares the most important characteristic quality of a base *i.e.*, it reacts with acids forming salt and water. They mutually *neutralise* each other.

$$\underset{\text{Sodium hydroxide}}{\mathrm{NaOH}\ (\text{alkali})} + \underset{\text{Hydrochloric acid}}{\mathrm{HCl}\ (\text{acid})} = \underset{\text{Sodium chloride}}{\mathrm{NaCl}\ (\text{salt})} + \underset{\text{water}}{\mathrm{H_2O}}$$

Expt. Take a little of caustic soda in a beaker, add sufficient water to it so that the solid dissolves on stirring. Add a drop or two of red litmus to the cooled solution, when the solution turns blue. Now add dilute solution of an acid (hydrochloric or sulphuric) to it, a little at a time, and stir after each addition. A time will come when the blue colour disappears and the whole solution turns red.

5-3. Salts :

By the term 'salt' we generally mean the familiar substance, sodium chloride, that we take everyday with our food. But in chemistry it has much wider meaning, it refers to a class of compounds. A *salt* is a compound that is formed by the replacement (partly or wholly) of the replaceable hydrogen atoms present in the molecule of an acid by a metal or a group of atoms acting like a metal, Potassium chloride (KCl), for example, is a salt, as it may be considered to have been formed by the displacement of hydrogen of hydrochloric acid by the metal potassium.

A salt consists of two parts—one or more atoms of metal (or group of atoms acting like a metal) constitute the first part and is known as basic part. The remainder is known as acidic part.

In aqueous solutions and in a state of fusion they are good conductors of electricity.

Questions

1. Write short notes on : acids, bases, alkalis and salts.

2. What is meant by an acid ? What is a base ? Is there any difference between a base and an alkali ? Give your answer with exampies.

3. Write in brief the characteristic qualities of 'acids' and 'alkalis'.

4. 'All alkalis are bases but all basses are not alkalis'—Discuss. Explain with suitable example the meaning of the term 'neutralisation'.

Sample questions for objective type tests

A. Alternate response type

Yes or No type —

(i) Is ammonia a base ? —

(ii) Does an acid contain one or more replaceable hydrogen atoms ? —

(iii) Are all bases alkalis ?

B. Recall type —

(i) An acid turns blue litmus —. —

(ii) Water soluble hydroxides are called —.

C. Completion type

Hydrogen atoms of acids may be — (a) or wholly — (a)
— (b) by certain metals or — (c) acting like a metal. — (b), — (c)
As a result metals or — (d) unite with the — (d)
acid radicals and new compounds known
as — (e) are formed. — (e)

CHAPTER SIX

Oxidation-Reduction

There are various types of chemical reactions. Nevertheless, there is a very important type that deserves special consideration and this is described as *oxidation-reduction.*

6·1. Oxidation-Reduction :

When carbon burns in air, carbon dioxide in formed. In this reaction oxygen combines with carbon. When copper oxide is heated in a current of hydrogen, metallic copper is obtained. Because of the romoval of oxygen from copper oxide. the latter turns into metallic copper. Originally the gain or addition of oxygen was termed as *oxidation* and the reverse process of this *i.e.*, the loss or removal of oxygen was termed as reduction.

Hence, **Oxidation**=*Gain or addition of oxygen,*

Reduction=*Loss or removal of oxygen.*

In the cases as illustrated above carbon is oxidised to carbon dioxide and the copper oxide reduced to copper.

Hydrogen, it is to be noted, is the exact 'chemical opposite' of oxygen ; they have a strong affinity for each other arising from their opposite chemical characters. The meaning of the terms has been extended to include 'addition' and 'removal' of hydrogen. Since the effect of addition of hydrogen is opposite to that of addition of oxygen.

Oxidation=*Removal of hydrogen,*

Reduction=*Addition of hydrogen.*

Later on chemists recognized that in behaviour oxygen is the emblem of electronegative elements while hydrogen is the emblem of electropositive elements. Hence, the scope of the terms has still further been extended and now we can say from the above that *oxidation is a process which involves an addition (or an increase in the proportion) of oxygen or any electronegative element or radical to a substance* or *a removal (or a decrease in the proportion) of hydrogen or any electro-positive element or radical from a substance.*

Sulphur, phosphorus, carbon, magnesium, etc. are converted into their oxides, when burnt. *e.g.*, $S+O_2=SO_2$. Here sulphur gets oxidised, as it gains oxygen.

Stannous chloride on reaction with chlorine is oxidised into stannic chloride ; it involves an addition of electronegative element chlorine. $SnCl_2+Cl_2=SnCl_4$.

Hydrogen sulphide, if passed into an aqueous solution of bromine, the former is oxidised to sulphur. $H_2S+Br_2=2HBr+S$. Hydrogen sulphide is relieved of its hydrogen and is thus oxidised.

Ozone oxidises potassium iodide to iodine. $O_3+2KI+H_2O=2KOH+I_2+O_2$. Oxidation has resulted from the removal of electropositive constituent potassium from potassium iodide.

The reverse process to that of oxidation being reduction, we now understand that *reduction is a process which involves an addition (or an increase in the proportion) of hydrogen or any electropositive element or radical to a substance* or *a removal (or a decrease in the proportion) of oxygen or any electronegative element or radical from a substance.*

If hydrogen sulphide is passed through chlorine water, chlorine is reduced to hydrogen chloride by gaining hydrogen.

$$Cl_2+H_2S=2HCl+S.$$

Mercuric chloride is reduced to mercurous chloride by addition of mercury (electropositive element), when the first and the last named substances are heated together, $HgCl_2+Hg=Hg_2Cl_2$.

Hydrogen reduces heated lead oxide to metallic lead by relieving the compound of its oxygen. $PbO+H_2=Pb+H_2O$.

Heated sodium reduces aluminium chloride to metallic aluminium by removing its electronegative constituent chlorine.

$$AlCl_3+3Na=3NaCl+Al.$$

6·2. Oxidising and reducing agents ;

Recollect the reaction between bromine and hydrogen sulphide. Here hydrogen sulphide being relieved of its hydrogen is *oxidised* to sulphur, while by gaining hydrogen the bromine is *reduced* at the same time to hydrogen bromide. It is thns evident that oxidation of one substance always necessitates the reduction of another and vice versa. The two processes, therefore, go hand in hand *i,e.*, occur simultaneously, In the example under consideration bromine is regarded as the oxidising agent. The substance that brings about oxidation and is itself reduced is termed as an *oxidising agent*. The opposite behaviour to that of bromine is exhibited by hydrogen sulphide, which, therefore, acts here as a reducing agent. The substance that brings about reduction and is itself oxidised is referred to as a *reducing agent.*

Questions

1. Explain what you mean by 'oxidation' and 'reduction'. Cite different examples.

2. 'Oxidation and reduction occur simultaneously'—Discuss.

3. Explain the terms : an oxidising agent and a reducing agent. Give one example in each case.

4. Classify the following changes as oxidation or reduction :

(i) $FeCl_3 \rightarrow FeCl_2$, (ii) $CuO \rightarrow Cu_2O$, (iii) $Al \rightarrow AlCl_3$? Adduce reasons in support of your statement,

Sample questions for objective type tests

A. Alternate response type

(i) *True or False type*

(a) The two processes oxidation and reduction are complementary. —

(b) Oxidation always means an addition of oxygen. —

(ii) *Yes or No type.*

(a) Does an oxidant in oxidation-reduction undergo reduction ? —

(b) Does the increase in the proportion of electro-positive constituent in a substance mean oxidation ? —

(c) Does a reductant in oxidation-reduction undergo oxidation ? —

B. Recall type

(i) The — process to that of oxidation is called reduction. —

(ii) If the proportion of electro-negative constituent in a substance is — the substance is said to be oxidised. —

CHAPTER SEVEN

Air

We are surrounded by an envelope of air, which we call atmosphere. We are scarcely aware of this colourless gaseous mixture, the existence of which we can perceive when the wind blows.

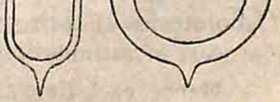
7-1. Composition of air :

Air is a mixture, the principal constituents of which are oxygen and nitrogen. Air freed from moisture and carbon dioxide contains nitrogen and oxygen approximately in the proportion of 4 : 1 by volume. It being a mixture, the composition of samples of air drawn from different localities and at different times is not quite constant. The average composition, however, in percentage by volume is shown below :

Nitrogen—78·1. Oxygen—20·90, Inert gases—0·94, Carbon dioxide—0·03.

Besides, water vapour, nitrous and nitric acid vapour, ammonia, ozone, sulphur dioxide, dust particles, bacteria, etc. are present in air. Sulphur dioxide is present in the air of industrial towns and ammonia in that of the dirty outskirts of cities.

Liquid air : As a result of investigations it is now clear that a gas must be sufficiently cooled, before it could be liquified. Moreover, for each gas there is a temperature, above which the pressure, however high, may be applied, the gas refuses to be liquified and shows no sign of liquifying. This being realized, the liquifaction of air presents no special difficulty. It is first made free from water vapour and carbon dioxide by treatment with fused calcium chloride and slaked lime respectively. The air thus purified is then liquified by subjecting it to high pressure and low temperature.



Dewar flask

Liquid air is kept in a double walled vessel known as Dewar flask, a vacuum being maintained between the walls. Both the walls—the outside surface of the inner and the inside of the outer one—are silvered. The transmission of heat is minimised and thus in this arrangement the sum total of heat transfer is very small.

The boiling point of liquid air is about −190°C. The presence of particles of ice and solid carbon dioxide makes the liquid look turbid. The filtered liquid is transparent and appears pale blue in colour due to the presence of liquid oxygen. The liquid evaporates, when poured in air ; and dense white clouds of condensed water vapours are produced. Under the influence of intense cold of liquid air many a substance acquires unusual properties. Fruits, raw meat, etc. for examples, become so hard as can be reduced to powder in a mortar. Flowers so treated can be crushed in hands. Mercury, as you know, is a liquid metal and changes under the influence of intense cold of liquid air into a malleable solid. Lead is not elastic but acquires this property on being cooled in liquid air and a lead bell emits a pure ringing sound, when struck.

Liquid air is a source of rare gases, which may be isolated from the former by fractional distillation. One of the industrial methods for obtaining oxygen and nitrogen involves fractionation of liquid air.

7-2. Nitrogen cycle :

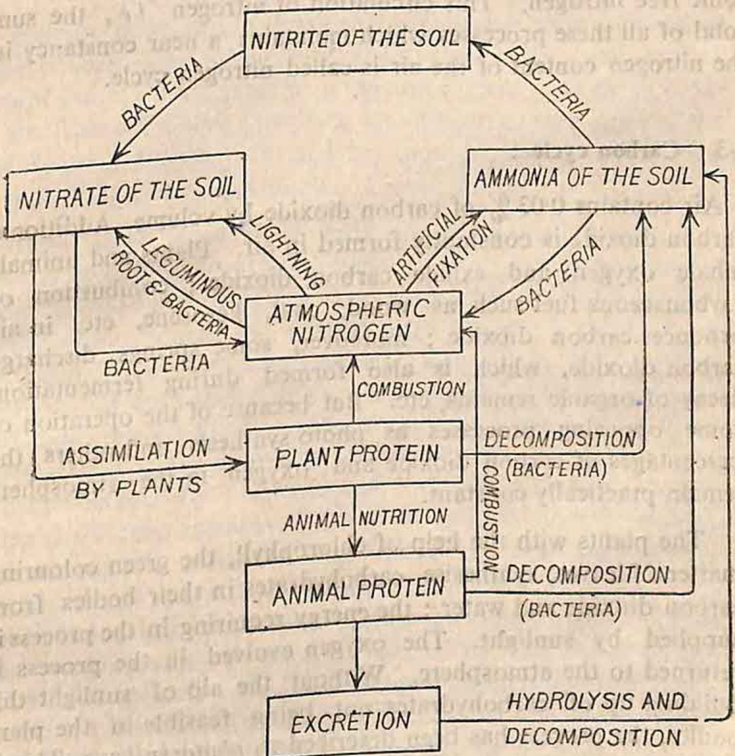
Nitrogen is essential to plant and animal life. Protein, a complex nitrogenous substance, is an important constituent of their tissues. Save and except a few bacteria, algae, fungi, etc, no plant or animal can directly assimilate nitrogen from air, which contains a large excess of it.

Through the action of lightning, nitrogen and oxygen of the air combine to form nitric oxide, which undergoes further oxidation and the product ultimately produces a certain amount of nitric acid on being dissolved in water. This is washed down by rains in the soil and there reacts with various bases, when soluble nitrates are produced. The plants absorb dissolved nitrates by the roots and synthesise the protein-foods in the laboratories of their own bodies. This is the most important process of assimilation of nitrogen of the air by plants.

Hence, each time a crop is taken from a given soil a depletion of nitrogen previously contained in the soil occurs. To make good the loss of soil nitrogen to some extent, paddy and the leguminous plants such as beans, peas, etc. may be grown in the same soil by turns. This is known as *crop rotation.* Some bacteria inhabit the roots of the leguminous plants causing their characteristic swellings, known as nodules. Bacteria thus living at the colonies in the roots convert free or atmospheric nitrogen present in the soil into nitrate, which is then assimilated by the plant.

Now-a-days the loss of nitrogen of the soil is replenished by administering the natural and especially the artificial fertilizers and thus eventually preventing the soil from becoming barren. Plants cannot utilise ammonia directly which, when added in the form of its compounds as fertilizers, is oxidised by some soil bacteria first to nitrites and then to nitrates to be assimilated by plants.

Animals, on the other hand, are ultimately dependent on plants for all of their nitrogen. Herbivorous animals feed upon plants, while the proteins of the former are utilised in turn by carnivora. We receive our necessary proteins from plants and animals.



Nitrogen cycle

This continuous removal of nitrogen should have long since resulted in the complete exhaustion of nitrogen content of the air, if there were not some mechanisms for replenishing the nitrogen losses. These mechanisms include :

Animal excretes nitrogenous compounds such as urea. Urea undergoes easy hydrolysis to ammonia and carbon dioxide. $CO(NH_2)_2 + H_2O \rightarrow 2NH_3 + CO_2$. Ammonia is also formed due to decomposition of some animal waste products. Death and consequential decay of plants and animals result in the formation of ammonia. This substance passes on to nitrates by the combined action of 'nitrosifying' and 'nitrifying' bacteria. Part of nitrates in the soil is utilised by plants and a part returned to the atmosphere as free nitrogen by the action of certain other soil bacteria, called 'denitrifying' bacteria. Moreover, combustion of plant and animal bodies results in the formation of some free nitrogen. This circulation of nitrogen *i.e.*, the sum total of all these processes which maintains a near constancy in the nitrogen content of the air is called nitrogen cycle.

7-3. Carbon cycle :

Air contains 0·03% of carbon dioxide by volume. Additional carbon dioxide is constantly formed in air. Plants and animals inhale oxygen and exhale carbon dioxide ; combustion of carbonaceous fuel such as wood, coal, kerosene, etc. in air produces carbon dioxide ; moreover, some springs discharge carbon dioxide, which is also formed during fermentation, decay of organic remains, etc. But because of the operation of some opposing processes as photo-synthesis and others the percentages of carbon dioxide and oxygen in the atmosphere remain practically constant.

The plants with the help of chlorophyll, the green colouring matter of leaves, synthesise carbohydrates in their bodies from carbon dioxide and water ; the energy requiring in the process is supplied by sunlight. The oxygen evolved in the process is returned to the atmosphere. Without the aid of sunlight this building up of carbohydrates not being feasible in the plant bodies, the process has been described as *photosynthesis*. This is also named as *carbon assimilation*. The total oxygen intake by plants during respiration is less than what they give over during carbon assimilation. Similarly the total carbon dioxide intake by plants during carbon assimilation is greater than what they give over during respiration. The plant, therefore, on balance, supplies oxygen to and removes carbon dioxide from the atmosphere.

Besides, some rocks also absorb carbon dioxide.

This natural arrangement of formation, removal and restoration through which a continuous circulation of carbon dioxide is maintained and its percentage in air is kept at nearly a constant value (0·03%) is called *carbon cycle.*

7-4. Inert gases :

Helium, neon, argon, krypton, xenon and radon are called *noble or inert gases*, because of their chemically inert nature. Except radon all are present in the atmosphere. As these occur in small amounts only, they are also known as *rare gases* of the atmosphere. Radon is formed by spontaneous decomposition of radium and hence the name.

These inert gases have won great practical importance. Because of lightness and non-inflammability, helium is used for filling balloons and dirigibles. A mixture of helium and oxygen, instead of air, is used for breathing by divers. It finds use in gas thermometers and creating inert atmosphere for welding active metals and their alloys. The electric filament lamps used today are gas-filled. A little nitrogen or argon is introduced into the bulb at low pressure. These lamps give brighter light than that emitted by a vacuum one. Besides efficiency, the life of the lamp also increases. Argon is also used to provide inert atmosphere in arc welding. Some of you are perhaps acquainted with advertisement signs displayed in towns and cities. These are popularly known as 'neon signs'. These decorative lights are, in fact, produced by electric discharge through various inert gases. Different gases emit different colours. Argon-filled lamp emits a blue light pleasing to the eye ; while lovely red sign is obtained from neon tube. Commercial krypton is used to produce the effect of daylight. Different and handsome colour signs are obtained by using neon in mixture with argon, mercury vapours, etc. Neon lights being visible in fog and mist act as a beacon to pilots.

Questions

1. Discuss only the bare outline of liquifying air and state the chief properties of liquid air.

2. What is carbon assimilation? Describe in brief the 'carbon cycle' of nature.

3. 'A world famine is imminent due to exhaustion of nitrogen'—Show that this prediction is baseless.

4. Describe the 'nitrogen cycle' in detail.

5. Mention the names of inert gases of the atmosphere and state their important uses.

Sample questions for objective type tests

A. Alternate response type

Yes or No type.

(a) Can plants synthesise the nitrogenous food in their own bodies? —

(b) Is radon named after radium? —

(c) Does the nitrogen content of the atmosphere gradually tend to decrease? —

B. Recall type

(i) A mixture of oxygen and —, instead of air, is used for breathing by divers. —

(ii) Some nitrogen and oxygen of air combine to form nitric oxide through the action of— in the atmosphere. —

(iii) Liquid air can be kept in double walled — flasks. —

C. Completion type

The plants with the help of—(a), the green colouring matter in leaves synthesise—(b), in their bodies from carbon dioxide and water; the energy requiring in the process is supplied by—(c). The—(d) evolved in the process is returned to the atmosphere.

—(a)
—(b)
—(c), —(d)

CHAPTER EIGHT

Commonly Known Gases

8-1. Oxygen

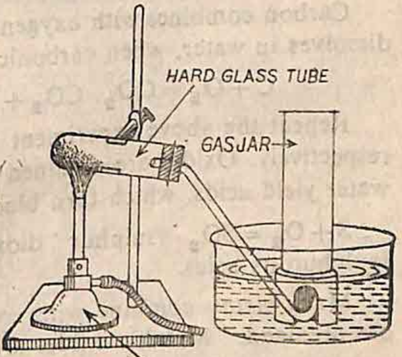
Symbol O, formulla O_2, atomic weight 16·000 (by definition).

Free oxygen occurs in the atmosphere. As compounds it is present in water, rocks, plants, etc. It is the most abundant of all the elements present in the earth's crust. The name *oxygen* was given to it by the French chemist Lavoisier.

Preparation : In the *laboratory* oxygen is usually obtained by heating a mixture of potassium chlorate and manganese dioxide.

$$2KClO_3[+MnO_2]=2KCl[+MnO_2]+3O_2.$$

Manganese dioxide facilitates the decomposition of potassium chlorate, but it itself remains unchanged in mass and composition and hence acts as a catalyst in the reaction. A hard glass test tube is half-filled with an intimate mixture of potassium chlorate and manganese dioxide (about 10—20%) and is then clamped as shown in the figure with its mouth end slightly inclined downwards. A delivery tube is fitted up through the cork at its mouth. The end of the delivery tube is dipped into water in a pneumatic trough. The evolved gas is eventually collected in gas jars by displacement of water.



Preparation of oxygen in the laboratory

There are many compounds which give oxygen, when heated. Mercuric oxide, for example, when heated, decomposes into metallic mercury and oxygen : $2HgO=2Hg+O_2$. Many nitrates decompose when heated. Lead nitrate, when heated, lead monoxide is formed, nitrogen dioxide and oxygen are evolved.

$$2Pb(NO_3)_2=2PbO+4NO_2+O_2.$$

A very convenient method of obtaining oxygen is to add water dropwise from a tap funnel on to sodium peroxide in a flask. No heating is required. $2Na_2O_2+2H_2O=4NaOH+O_2$.

Properties : *Physical*—it is a colourless, tasteless and odourless gas. It is slightly heavier than air and is slightly soluble in water.

Chemical—It is non-combustible but a good supporter of combustion.

Expt. Introduce the glowing end of a wood splint into a jar of oxygen, The splint is rekindled, but the gas does not burn.

Oxygen, in fact, is highly reactive. Some *non-metals* burn in oxygen and oxides are formed.

Expt. Heat a piece of charcoal in a deflagrating spoon and now introduce the glowing charcoal, while on spoon, into a jar of oxygen. Charcoal begins to burn with a flame. Pour some water into the jar, shake well, test the solution with a blue litmus paper, the latter turns red.

Carbon combines with oxygen forming carbon dioxide, which dissolves in water, when carbonic acid is obtained.

$$C+O_2=CO_2,\ CO_2+H_2O=H_2CO_3$$

Repeat the above experiment with phosphorus and sulphur respectively. Oxides are obtained and these on combination with water yield acids, which turn blue litmus red.*

$S+O_2=SO_2$ (sulphur dioxide) ; $H_2O+SO_2=H_2SO_3$ (sulphurous acid).

Many *metals* combine with oxygen, when heated and oxides are obtained. Metallic oxides are usually basic in nature.

Sodium burns brilliantly emitting a yellow light. The oxide (mainly peroxide) formed reacts with water and the resulting solution, which contains sodium hydroxide, turns red litmus blue.* $2Na+O_2=Na_2O_2$, $2Na_2O_2+2H_2O=4NaOH+O_2$. Potassium, however, burns with a violet flame instead of a yellow one. A piece of burning magnesium ribbon, if lowered into a jar of oxygen, the metal continues to burn with a dazzling white flame. The product magnesium oxide dissolves slightly in water forming magnesium hydroxide and a red colour is produced, when phenolphthalein is added to it.

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* See the chapter on Acids, Bases and Salts.

Oxygen reacts with *some compounds* at room temperature. Nitric oxide, on coming in contact with oxygen, yields brown nitrogen dioxide. $2NO+O_2=2NO_2$. On reaction with oxygen sulphurous acid is converted into sulphuric acid and ferrous salt into ferric salt. $2H_2SO_3+O_2=2H_2SO_4$.

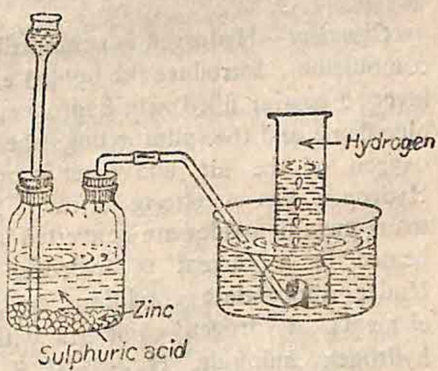
An acid solution of chromous chloride, alkaline potassium pyrogallate may be used as absorbents for the gas.

8-2. Hydrogen

Symbol H, formula H_2, atomic weight 1·008.

In free state hydrogen does not occur in nature to any appreciable (1 part per million) extent. But large volumes are present in the outer atmosphere of the sun. Compounds of hydrogen are numerous, abundant and widely distributed in nature. Water contains one-ninth of its weight of hydrogen. Hydrogen is a common constituent of mineral oils and most animal and vegetable products. Lavoisier named it hydrogen *i.e.*, water former or water-producer from Greek *hudor*-water and *genon*—to produce.

Preparation : *Laboratory method* : Zinc granules are taken in a two-necked Woulfe's bottle fitted with a thistle funnel and a delivery tube as shown in the figure. Enough water is added to cover the zinc, care being taken so that the end of the thistle funnel dips under the liquid. The arrangement is made airtight. Now conc. sulphuric acid, one volume of acid for five volumes of water, is added down the funnel. Brisk effervescence starts, the gas is first allowed to escape, it sweeps out all the air present and eventually the evolved hydogen is collected in gas jars by displacement of water. Zinc dissolves by reacting with acid. when hydrogen evolves.



Preparation of hydrogen in the laboratory

$$Zn+H_2SO_4=ZnSO_4+H_2.$$

A Kipp's apparatus may conveniently be used for preparing the gas. *Zinc* is taken in the middle globe and dilute acid is poured into the top one.

Water is the cheapest source of hydrogen. Some metals decompose water under certain conditions dependent on the nature of the metal. Metals like sodium, potassium etc. decompose water at *ordinary temperature*, but magnesium powder or zinc copper couple decomposes *hot* water, while *red hot* iron, Zinc etc. decompose *steam*. In all the cases, hydrogen is liberated and the reactions occurring are : $2Na+2H_2O=2NaOH$ (sodium hydroxide)$+H_2$; $Zn+2H_2O=Zn(OH)_2+H_2$; $3Fe+4H_2O$ (steam)$\rightleftharpoons Fe_3O_4$ (ferroso-ferric oxide)$+4H_2$.

Alkalis contain hydrogen. Solution of strong alkali like NaOH or KOH, if heated with zinc or aluminium, hydrogen is liberated. $Zn+2KOH=K_2ZnO_2$ (potassium zincate)$+H_2$.

Properties : *Physical* : Hydrogen is a colourless, odourless and tasteless gas. It is non-poisonous, though unable to support respiration. It is practically insoluble in water. It is the lightest gas and air is 14·4 times as heavy as hydrogen. Balloons filled with hydrogen float in air.

Chemical—Hydrogen is combustible but not a supporter of combustion. Introduce the lighted end of a wood splint into an inverted gas-jar filled with hydrogen, the gas burns with a pale blue flame and the splint extinguishes. Hydrogen combines with oxygen of the air and water is produced, $2H_2+O_2=2H_2O$. Hydrogen has a strong affinity for oxygen and that is why many metallic oxides are converted into respective metals, when heated in a current of hydrogen (see oxidation-reduction). Under appropriate conditions hydrogen combines with many elements. Hydrogen combines with boiling sulphur forming hydrogen sulphide, $H_2+S=H_2S$. Hydrogen and nitrogen combine directly but slowly on sparking. Under high pressure and temperature and in presence of iron powder (accompanied by 'promoters') as catalyst, hydrogen and nitrogen combine to form ammonia. $3H_2+N_2\rightleftharpoons 2NH_3$. In presence of light it reacts with chlorine producing hydrogen chloride. $H_2+Cl_2=2HCl$. It combines with only a few metals like sodium, calcium etc. to form hydrides, $2Na+H_2=2NaH$.

Nascent hydrogen : An element in said to be in the nascent state at the moment of its generation from a compound. Nascent hydrogen is more active than ordinary hydrogen.

Pass hydrogen gas through a little of potassium permanganate solution acidified with dil. sulphuric acid,—no change in colour of the solution occurs. Again add a few pieces of zinc to another portion of the acidified solution.—The nascent *i.e.*, just born hydrogen decolourises permanganate by reduction, $2KMnO_4+8H_2SO_4+10H=K_2SO_4+2MnSO_4+8H_2O$. The same effect will be observed with ferric chloride, which is reduced by nascent hydrogen and not by molecular hydrogen to colourless ferrous chloride:

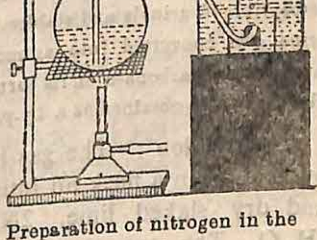
$$FeCl_3+H=FeCl_2+HCl.$$

8-3. Nitrogen

Symbol N, Formula N_2, atomic weight 14.

Elementary nitrogen occurs in the atmosphere, it constitutes about four-fifth of the total volume of air, combined nitrogen in the form of proteins is a constituent of plant and animal tissues. As sodium nitrate in Chile and as nitre (KNO_3) in soils of some places significant quantities of it are available. It being a non-supporter of respiration Lavoisier named it *azote* meaning *no life*. The present name nitrogen *i.e*, nitre-producer originates from the fact that it is a constituent of nitre.

Preparation : *Laboratory method*—A concentrated solution containing equimolecular quantities of ammonium chloride and sodium nitrite is gently and carefully heated in a flask fitted with a thistle funnel and a delivery tube, which leads into a trough of water. At first air of the flask is allowed to escape and then the gas is collected in gas-jars by displacement of water. The gas may be dried by conc. H_2SO_4. $NH_4Cl+NaNO_2=NH_4NO_2+NaCl$; $NH_4NO_2=N_2+2H_2O$.



Preparation of nitrogen in the laboratory

Properties : It is a colourless, odourless and tasteless gas. It is almost as heavy as air and is very slightly soluble in water.

Nitrogen is noted for its inertness. Under the usual laboratory conditions it does not react with other elements. It is neither combustible nor a supporter of combustion. Certain metals such as calcium, magnesium etc. combine at high temperatures forming nitrides. $3Ca+N_2=Ca_3N_2$; $3Mg+N_2=Mg_3N_2$. These nitrides react with water to yield ammonia, $Mg_3N_2+6H_2O=2NH_3+3Mg(OH)_2$. In a mixture with oxygen electric sparks make them combine slowly to nitric oxide. $N_2+O_2\rightleftharpoons 2NO$. Under suitable conditions (p. 50) nitrogen and hydrogen combine together to form ammonia.

Some simple reactions leading to isolation of nitrogen from air :

You know that air is a mixture. If air freed from moisture and carbon dioxide be allowed to pass over red hot *copper*, oxygen of the air combines with copper, while nitrogen remains unchanged. $2Cu+O_2=2CuO$.

If *mercury* be heated or *phosphorus* be burnt in air, oxygen combines with the element concerned and nitrogen of the air remains unaffected, $2Hg+O_2=2HgO$, $4P+5O_2=2P_2O_5$.

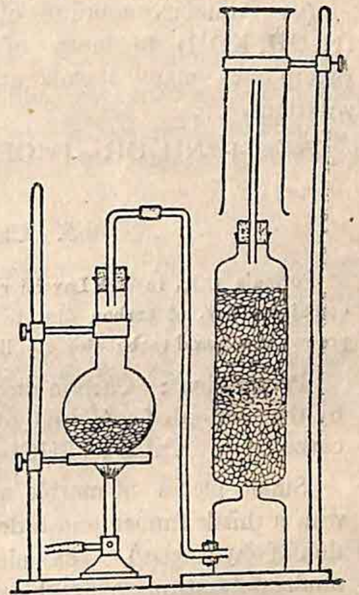
Nitrogen may also be isolated by absorbing oxygen of the air in a suitable *absorbent*.

8-4. Ammonia

Formula NH_3, molecular weight 17.

Ammonia occurs only in traces in the atmosphere. It is formed by the decomposition of nitrogenous organic matter. Sharp smell of ammonia is perceptible in urinals and stables. Urea an organic compound present in the urine gets converted into ammonium carbonate by the action of bacteria. Ammonium carbonate in its turn undergoes decomposition to give ammonia. The gas is also obtained as a by-product in the distillation of coal.

Preparation : The gas is usually prepared in the *laboratory* by heating in a flask an intimate mixture of ammonium chloride and dry slaked lime. $2NH_4Cl+Ca(OH)_2=2NH_3+CaCl_2+2H_2O$. The gas that evolves is dried by passing through a tower filled with small lumps of quicklime, which absorbs the accompanying water. The dried gas is eventually collected in gas-jars by downward displacement of air.


Preparation of ammonia in the laboratory

Ammonia is a base ; with conc. sulphuric acid it forms ammonium sulphate. Unless very pure and almost completely dry, the gas reacts with phosphorus pentoxide. Calcium chloride absorbs the gas forming $CaCl_2$, $8NH_3$. That is why none of these reagents can be utilised for drying the gas.

Properties : *Physical*—Ammonia is a colourless gas with a sharp pungent smell. It can be easily liquefied by cold or pressure. It is lighter than air and highly soluble in water. The aqueous solution reacts alkaline, it turns red litmus blue. The concentrated aqueous solution is known as liquor ammonia.

Chemical—(*i*) It is neither combustible in air nor a supporter of combustion. It burns with a greenish yellow flame in an atmosphere of oxygen. A mixture of ammonia and oxygen is explosive, feeble explosion occurs, when it is kindled.

$$4NH_3+3O_2=6H_2O+2N_2.$$

(ii) Heated alkali metals and barium decompose ammonia, when an amide is formed and hydrogen is liberated.

$2NH_3+2Na=2NaNH_2+H_2$. This proves that ammonia contains hydrogen.

(iii) Many heated oxides of metal such as PbO, CuO are reduced to metallic condition. $3CuO+2NH_3=3Cu+N_2+3H_2O$. That ammonia contains nitrogen may be proved by this reaction.

(iv) Ammonia is a basic substance. It reacts with acids yielding ammonium salts. $NH_3+HCl=NH_4Cl$ (ammonium chloride).

(v) Aqueous solution of ammonia resemble typical alkalis (NaOH, KOH) in many of its properties. Aqueous solution precipitates many metals as their hydroxides from their salt solutions.

$$FeCl_3+3NH_4OH=Fe(OH)_3\text{(ferric hydroxide)}\downarrow+3NH_4Cl.$$

8-5. Carbon dioxide.

Formula CO_2, in 1788 Lavoisier proved that it is an oxide of carbon. Air contains 0·03% of carbon dioxide by volume. Combined carbon dioxide is present in mineral carbonates, *e.g.* limestone, magnesite, dolomite, etc.

Preparation : Carbon dioxide is prepared in the *laboratory* by the action of cold and dil. hydrochloric acid on calcium carbonate : $CaCO_3+2HCl=CaCl_2+H_2O+CO_2$.

Small pieces of marble are taken in a Woulfe's bottle fitted with a thistle funnel and a delivery tube (just as in the preparation of hydrogen,). The pieces are covered with water and moderately strong hydrochloric acid is poured down the funnel. As the acid comes in contact with the marble, brisk effervescence occurs due to evolution of the gas, which is collected by the upward displacement of air.

Sulphuric acid, instead of hydrochloric acid, cannot be used in the method, since the marble soon gets coated with a layer of rather insoluble calcium sulphate, which prevents the marble from coming in contact with the acid and consequently the reaction ceases.

As the reaction requires no heating, it may conveniently be carried out by charging a Kipp with marble and dil. hydrochloric acid for ready supply of the gas, whenever required.

Properties: *Physical* : (i) Carbon dioxide is a colourless gas with a weakly acid taste and faintly pungent smell. (ii) It is 1·5 times as heavy as air. It can be poured on a flame to put it out. (iii) It does not support respiration. (iv) It can be easily liquefied (below 31°) by compression. When the liquid is allowed to evaporate rapidly, a portion of it solidifies ; the solid is known as *dry ice*. (v) It is fairly soluble in water and the solubility increases with the rise of pressure. Aerated waters are charged with the gas under pressure.

Chemical : (i) It is an acidic oxide, aqueous solution of it acts as an acid and turns blue litmus dull red. With water it forms an unstable and weak acid known as carbonic acid.

$$CO_2+H_2O \rightleftharpoons H_2CO_3.$$

It forms carbonates or bicarbonates on reaction with bases and alkalis. $CaO+CO_2 \rightleftharpoons CaCO_3$; $2NaOH+CO_2=Na_2CO_3+H_2O$, and with excess of the gas sodium carbonate passes on to sparingly soluble bicarbonate :

$$Na_2CO_3+H_2O+CO_2 \rightleftharpoons 2NaHCO_3.$$

When carbon dioxide is allowed to pass through lime water, it first forms calcium carbonate, (insoluble) and then with excess of the gas a bicarbonate, which is soluble. That is why the lime water first turns milky and the milkiness eventually disappears. Carbon dioxide evolves on boiling the clear liquid, which turns turbid again.

$$Ca(OH)_2+CO_2=CaCO_3+H_2O.$$
$$CaCO_3+H_2O+CO_2 \rightleftharpoons Ca(HCO_3)_2.$$

(ii) Carbon dioxide is neither combustible nor a suporter of combustion. But burning magnesium, sodium and potassium continue to burn in the gas, when pure carbon separates out. $2Mg+CO_2=2MgO+C$, $4K+3CO_2=2K_2CO_3+C$. This reaction shows that carbon dioxide contains carbon.

(ii) Red hot carbon, zinc or iron powder may reduce it to carbon monoxide, $CO_2+C=2CO$.

8·6. Hydrogen sulphide, H_2S

Traces of hydrogen sulphide occur in the atmosphere. Silver article tarnishes in air due to formation of a layer of silver sulphide, which is black. It also occurs in volcanic gases and in the water of certain springs. Additional hydrogen sulphide is being poured into the atmosphere from decay or decomposition of sulphur-containing materials. The foul odour of rotten eggs or putrefied hides is partly due to hydrogen sulphide.

Preparation : It is usually prepared in the *laboratory* by the action of dil. sulphuric or better dil. hydrochloric acid on ferrous sulphide taken in a Woulfe's bottle fitted (just as in the case of hydrogen) with a thistle funnel and a delivery tube. As soon as the small pieces of ferrous sulphide come in contact with the acid, the reaction occurs and the evolved gas is collected by up-ward displacement of air (as in the case of carbon dioxide) or by displacement of hot water. $FeS+2HCl=FeCl_2+H_2S$. The gas may be dried with phosphorus pentoxide.

Hydrogen sulphide, free from hydrogen may conveniently be obtained by the action of conc. hydrochloric acid on antimony sulphide ; $Sb_2S_3+6HCl=2SbCl_3+3H_2S$.

For intermittent but ready supply of the gas in the laboratory the reaction between ferrous sulphide and dilute sulphuric acid is carried out in a Kipp's apparatus.

$$FeS+H_2SO_4=FeSO_4+H_2S.$$

[Demonstration of the Kipp's and the explanation of its mechanism by teacher in the class will be helpful to the students for a proper understanding of the process.]

Properties : *Physical :* It is a colourless gas having the smell of rotten eggs. It is heavier than air. It is moderately soluble in cold water but insoluble in hot water. It is poisonous and if inhaled for a longer period it may cause death.

Chemical : The gas itself is combustible in air, but it does not support combustion. It burns at the jet with a blue flame yielding sulphur dioxide and water, $2H_2S+3O_2=2H_2O+2SO_2$; but in limited supply of air it burns to sulphur and water and hence in a gas jar sulphur deposits due to incomplete combustion,

$$2H_2S+O_2=2H_2O+2S.$$

H_2S-water is a weak acid, reaction with alkalis gives two types of salts.

$$2NaOH+H_2S=Na_2S+2H_2O,\ Na_2S+H_2S=2NaHS.$$

Hydrogen sulphide is a good reducing agent. Thus it reduces halogen in aqueous solution or suspension to halogen hydracids with separation of sulphur. $Cl_2+H_2S=2HCl+S$, $I_2+H_2S=2HI+S$. Moist sulphur dioxide, conc. sulphuric acid, nitric acid are reduced by hydrogen sulphide. $SO_2+2H_2S=2H_2O+3S$, $2HNO_3+H_2S=2H_2O+2NO_2+S$, $H_2SO_4+H_2S=2H_2O+SO_2+S$. It is for this reducing action neither nitric acid can be used for preparing the gas nor sulphuric acid for drying it. It decolourises solutions of ferric chloride and acidified potassium permanganate by reduction. $2FeCl_3+H_2S=2FeCl_2+2HCl+S$, $2KMnO_4+3H_2SO_4+5H_2S=K_2SO_4+2MnSO_4+8H_2O+5S$.

Some metallic sulphides with characteristic colours are precipitated by hydrogen sulphide under varying conditions from their salt-solutions, *e.g.*, $CuSO_4 + H_2S = CuS$ (black) $\downarrow + H_2SO_4$.

8-7. Sulphur dioxide, SO_2.

Priestly in 1774 obtained the gas by heating mercury with conc. sulphuric acid. It is present in the volcanic gases and in water of some springs. It also occurs in the air of industrial belts. Sulphur, sulphide minerals such as iron pyrites FeS_2, when burnt give sulphur dioxide. $S + O_2 = SO_2$, $4FeS_2 + 11O_2 = 2Fe_2O_3 + 8SO_2$.

Preparation : In the *laboratory* it is usually obtained by reducing conc. sulphuric acid with metallic copper. $Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2$.

Copper turnings and conc. sulphuric acid are heated together in a flask fitted with a thistle funnel and a delivery tube (cf. Nitrogen preparation). As soon as the gas begins to evolve, heating is stopped. The gas being soluble in water and heavier than air is collected by upward displacement of air. A purer specimen free from moisture is obtained by allowing the gas to pass through washer containing conc. sulphuric acid and collecting over mercury. Calcium chloride or phosphorus pentoxide may also be used in drying the gas.

The gas may, however, be conveniently made in the laboratory by the action of a strong acid on sodium hydrogen sulphite.

$NaHSO_3 + H_2SO_4 = NaHSO_4 + SO_2 + H_2O$, the reaction requires no heating.

Properties : *Physical* : Sulphur dioxide is a colourless gas with a suffocating odour. It is heavier than air and is soluble in water. It can be liquefied by cooling in a freezing mixture or by compression even at ordinary temperature.

Chemical : The gas is neither combustible nor a supporter of combustion. The gas, however, is converted into sulphur trioxide by catalytic (platinum, vanadium pentoxide, etc.) oxidation ; $2SO_2 + O_2 = 2SO_3$. Heated potassium, finely divided iron or tin burn in the gas. $4K + 3SO_2 = K_2SO_3 + K_2S_2O_3$, $3Fe + SO_2 = 2FeO + FeS$.

Aqueous solution of sulphur dioxide is acidic, as it forms a weak unstable acid with water ; it turns blue litmus red. $SO_2+H_2O \rightleftharpoons H_2SO_3$ (sulphurous acid). Sulphurous acid forms two types of salts with alkalis. With lime water calcium sulphite is first formed, and with excess of the gas the insoluble sulphite passes on to soluble bisulphite. That is why the lime water first turns turbid and then the turbidity disappears.

$$Ca(OH)_2+SO_2=CaSO_3+H_2O\ ;$$
$$CaSO_3+SO_2+H_2O=Ca(HSO_3)_2.$$

It is usually a reducing agent. Yellow ferric chloride, pink potassium permanganate solutions are decolourised by reduction, while orange-yellow acidified dichiomate solution turns green :

$$2FeCl_3+SO_2+2H_2O=2FeCl_2+2HCl+H_2SO_4\ ;$$
$$2KMnO_4+5SO_2+2H_2O=K_2SO_4+2MnSO_4+2H_2SO_4\ ;$$
$$K_2Cr_2O_7+3SO_2+H_2SO_4=K_2SO_4+Cr_2(SO_4)_3+H_2O.$$

Chlorine, bromine and iodine are reduced by sulphur dioxide to corresponding hydracids. $Cl_2+SO_2+2H_2O=2HCl+H_2SO_4$. It may act also as an oxidising agent as is seen when it reacts with hydrogen sulphide (p. 56).

Sulphur dioxide acts as a mild bleaching agent in presence of moisture and hence is used in bleaching delicate materials like silk, wool, etc., which are likely to be damaged by chlorine.

Questions

Oxygen

1. Describe the laboratory method of preparation of oxygen and state its chief properties.

2. A colourless, odourless gas evolves, when mercuric oxide is heated in a hard glass test tube. What is the name of the gas ? Give the equation. Describe the laboratory method of preparation of the gas.

3. State the role played by manganese dioxide in the preparation of oxygen from potassium chlorate. The word oxygen, means acid producer ; cite two examples to justify the name and two others that go against it.

4. Name three such substances as will give oxygen, when heated singly or mixed with others. Express the changes by equations only. Can oxygen be prepared without applying heat ? State the equation of the reaction involved in such a process.

5. Describe with equations what happens when—
   - (i) a jar containing nitric oxide is made open in air,
   - (ii) carbon, sulphur, sodium and magnesium are separately burnt in oxygen.

6. (a) How can you obtain oxygen from air (a) physically and (b) chemically? (Simply indicate in short).

[**Hints :** Fractionation of liquid air ; $Hg \xrightarrow{air} HgO \xrightarrow{heat} O_2$].

(b) Sulphur, sodium and hydrogen are separately burnt in air. Show that the resulting oxides differ chemically.

[**Hints :** Water does not change the colour of litmus.]

Hydrogen

1. Describe the laboratory method of preparation of hydrogen. State its principal properties.

2. Describe two chemical reactions for the preparation of hydrogen from water. How would you know that the collected gas is hydrogen? What is nascent hydrogen? How would you prove that nascent hydrogen is more active than the gaseous specimen?

3. Describe one reaction each for the preparation of hydrogen from acid, alkali solution and steam.

4. What happens, when,
   - (i) steam is passed over hot zinc,
   - (ii) hydrogen is passed over heated cupric oxide, and
   - (iii) solution of potassium hydroxide admixed with zinc be heated?

Nitrogen

1. Describe the laboratory method of preparation of nitrogen. Write in brief what you know about its properties.

2. Describe the main reaction involved in each of the processes known to you for isolation of nitrogen from air chemically. Why was the gas named as 'azote' and why is it now named as 'nitrogen'?

3. Under what condition does nitrogen react with each of the following : hydrogen, oxygen, calcium and magnesium? State the reactions with equations.

4. An inert gas is passed over heated magnesium, when the latter absorbs the gas. The product is treated with water when a new gas is obtained. The new gas if allowed to pass over heated copper oxide the original inert gas is recovered. What is the original gas? Explain the process stating with equations the reactions involved.

[**Hints :** $NH_3 \xrightarrow[CuO]{heated} N_2$, see ammonia. p. 53],

Ammonia

1. How is dry ammonia prepared in the laboratory? Why is it that none of the following reagents can be used for drying the gas: conc. sulphuric acid, phosphorus pentoxide and calcium chloride?

2. How and under what conditions does ammonia react with the following:

Copper oxide, sodium and oxygen? Give equations.

3. State with equations the reactions that go to prove that
   (a) ammonia may act as a reducing agent,
   (b) ammonia contains nitrogen, and
   (c) ammonia contains hydrogen.

Carbon dioxide

1. How is carbon dioxide prepared in the laboratory? Explain why hydrochloric acid and not sulphuric acid is used in the process? State the chief properties of carbon dioxide.

2. What is 'dry ice'?

'Effervescence occurs, when a soda-water bottle is opened up in air'—Explain.

State with equation a reaction which indicates that carbon dioxide contains carbon.

3. Describe with equations the changes that occur in the following processes:

(a) carbon dioxide is gradually passed through lime water and the latter subsequently boiled,

(b) carbon dioxide is allowed to pass over red-hot carbon, and

(c) a burning magnesium ribbon is introduced into a jar of carbon dioxide.

Hydrogen sulphide

1. How is hydrogen sulphide prepared in the laboratory? Explain why conc. sulphuric acid is not used for drying the gas.

2. State with equation a reaction that leads to formation of hydrogen sulphide free from hydrogen.

How can you get a ready and intermittent supply of the gas in the laboratory?

Explain why nitric acid is not used in the preparation of the gas.

3. State the changes that occur when hydrogen sulphide is separately passed through each of the following:—

(i) nitric acid, (ii) iodine suspended in water, (iii) conc. sulphuric acid and (iv) ferric chloride solution. Give equations.

Sulphur dioxide

1. How is sulphur dioxide prepared in the laboratory? How is it dried and a purer specimen obtained? State its principal properties.

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Madhyamik (Secondary) Examination, 1976

PHYSICAL SCIENCE

Group A

1. (a) Which is measured by a common balance—the weight of a body or its mass ? 2

(b) How will the weight of a body be changed if it is taken from the earth of the moon ? How will its mass be changed ?

(c) What do you understand by "Conservation of Mass" ? When coal is burnt, the ash that is left has mass much smaller than that of the coal. What happens to the missing mass ? 3

2. (a) "The mass of a hydrogen atom is almost equal to that of a proton." What is the justification of this statement ? 2

(b) What is an isotope ? Uranium has two isotopes whose atomic weights are 235 and 238. Their atomic number is 92. In what way do the structures of their nucleus differ and in what way do they agree ? 2+4

3. (a) State Charles' Law and Boyle's Law. 4

(b) Are the molecules of a gas always at rest ? If not, when does the speed of the molecules increase ? 2

(c) At a pressure of 76 cm. the volume of a gas is 500 c.c. If the temperature remains unchanged and if the pressure be changed to 38 cm. what will be the volume ? 2

Group B

4. (a) State Newton's first law of Motion. Define force. 4

(b) Why is it dangerous to jump out of a moving carriage with the face directed opposite to the direction of motion of the carriage ? 4

5. (a) Explain with the help of a diagram the total internal reflection. Give an example of total internal reflection in a natural phenomenon. 3

(b) How far from a convex lens should an object be placed so that its image is real. 2

(c) How will you show that white light is made up of lights of different colours ? What is the name of the band of colours so obtained ? 3

6. (a) For refraction at a plane glass surface from air draw a diagram showing the incident ray, the refracted ray, the angle of incidence and the angle of refraction. 4

(b) If the angle of incidence be 45°, what will be the angle of reflection ? 2

(c) Is the image seen in a plane mirror real or virtual ? 2

7. (a) We hear the sound of thunder a few seconds after we see the flash of lightning in the sky. Why ? 2

(b) A bell is ringing inside a bell jar. If the air be gradually pumped out from within the bell jar the sound becomes fainter and fainter. Why ? 2

(c) State two uses of ultrasonic sound. 4

8. (a) On the body of an electric bulb it is written '220 volts 40 watts'. What is understood from this ? 4

(b) When an electric bill is received in a house it is found that 24 units was spent last month. What is this unit of ? What is the name of this unit ? 2

(c) There are two copper wires of the same length but one is thicker than the other. If they are connected across the same battery, state in which wire the electric current flowing is greater. 2

9. (a) There is a straight wire below a magnetic needle and parallel to it. If a current be passed through the wire from south to north, in which direction east or west —will the North pole of the needle be deflected ? State the law by which you determine this. 2+4

(b) What kind of change do the cathode rays carry ? How can they be deflected from their path ? 2

Group C

10. (a) Define :—

(i) Solubility. (ii) Oxidation.

(b) How many atoms are there in a molecule of sulphuric acid ?

(c) What is the volume of the gas at standard temperature and pressure, that will be formed by the reaction of 50 gms. of calcium carbonate with excess of hydrochloric acid ? How many molecules are there in the gas generated ? 3+1+4

(Atomic weight of calcium=40)

11. (a) Balance the following equations :—

(i) $H_2SO_4+Cu=CuSO_4+SO_2+H_2O$

(ii) $NH_3+O_2 \quad =N_2+H_2O$

(b) Define covalency.

(c) State with reasons whether the following changes are physical or chemical :—

(i) Magnetisation of iron. 2+2+4

(ii) Electrolysis of water.

12. (a) Describe, giving a sketch of the apparatus used, the preparation of ammonia gas in the laboratory.

(b) Show with the help of two examples that it is basic in nature. 5+2+1

(c) Name a fertiliser derived from it.

Form No. 3.

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